chapter

14 Electrochemical Cells

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Since their invention in 1888, vehicles powered entirely by electricity have drifted in and out of fashion (**Figure 1**). Many experts predict that in the next decade, electric vehicles will finally make a breakthrough. Electric power is slowly becoming a viable alternative to gasoline power, thanks to a combination of political, economic, and environmental factors. One advantage of electric cars over gasoline-fuelled cars is that they produce less pollution. Also, while cars powered by gasoline engines are about 15% efficient, many electric cars are 90% efficient. (Of course, overall efficiency and environmental impacts depend on how the electricity and gasoline are produced.) Other attractive features of electric vehicles are that they are nearly silent and require minimal maintenance.

The biggest obstacle to the widespread use of electric cars is the lack of a powerful, lightweight, inexpensive battery. Scientists and engineers are researching alternatives to the common lead–acid battery. Perhaps the most promising alternative is a battery that runs continuously as fuel is supplied. One such alternative is the aluminium–air fuel cell, which uses aluminium metal as the fuel and oxygen from the air to produce electricity. Another possibility is a fuel cell in which a hydrogen-rich fuel and oxygen from the air produce electricity.

Redox reactions can produce electricity and, conversely, electricity can cause redox reactions. Many materials that we take for granted were virtually unknown until the process of electrolysis made their production possible. Aluminium, chlorine, hydrogen, sodium hydroxide, magnesium, and copper are produced in large quantities by electrolytic processes. In this chapter, you will learn how batteries are made, how electricity can be used to produce chemicals, and how science and technology work together in the development of electrochemical processes.

STARTING Points

Answer these questions as best you can with your current knowledge. Then, using the concepts and skills you have learned, you will revise your answers at the end of the chapter.

1. What concepts can we use to explain how electrochemical cells work?
2. Describe the relationship between science and technology in the development of electrochemical cells.
3. List the types and uses of a variety of common electric cells. Include an assessment of the impact of each one on our lives.

Career Connections:
Materials Engineering Technologist; Chemical Technologist

Figure 1
Today's electric cars are a viable alternative to gasoline-powered cars in urban environments. An electric car has electric motors that are powered by rechargeable batteries.

Exploration A Simple Electric Cell

A cell that produces electricity can be amazingly simple because it uses very common materials and requires no technical expertise to construct. Anyone can make one and then improve its efficiency without much understanding of the scientific principles involved. That is why the electric cell was used for more than 100 years before scientists understood how it worked.

Materials: copper and zinc metal strips (or any two different metals); steel wool; orange, apple, and potato (and other fruits or vegetables); LCD clock; voltmeter (or multimeter) with leads

- Clean the metal strips with steel wool to remove any coating or oxides.
- Stick both metal strips into the orange. Make sure that the metal strips are not in contact inside or outside the orange.
- Momentarily touch the leads (red—positive; black—negative) from the voltmeter, one to each metal strip. Now reverse the leads and test again.

(a) Record and describe what happened in each case.

- Connect the leads to the LCD clock, paying attention to positive and negative connections.

(b) Does the clock work? If it does not, suggest a solution to make it work. Try it.

(c) Explain, in your own words, what you think happened in (b).

- Repeat the process using other fruits and vegetables.

(d) Which fruit or vegetable seemed to be the best at producing electricity?

(e) What do all fruits and vegetables have in common?

(f) How could you improve upon your electric cell?

14.1 Technology of Cells and Batteries

Before 1800, scientists knew that static electricity was produced by the friction created by two moving objects in contact. They discovered ways of storing the charges temporarily, but when the energy was released in the form of an electrical spark, it could not be put to practical use. Practical applications of electricity were developed after 1800, the year in which Alessandro Volta announced his invention of the electric cell.

Volta invented the first electric cell but he got his inspiration from the work, almost 30 years earlier, of the Italian physician Luigi Galvani. Galvani noticed that the muscles in a frog's leg would twitch when a spark hit the leg. Galvani's crucial observation was that two different metals could make the muscle twitch. Unfortunately, Galvani thought his discovery was due to some mysterious "animal electricity." It was Volta who recognized that this effect had nothing to do with animals or muscle tissue, and everything to do with conductors and electrolytes, as you observed in the Exploration at the beginning of this chapter.

Cells and Batteries

Although an *electric cell* is a device that continuously converts chemical energy into electrical energy, the electric cells that Volta invented produced very little electricity. Eventually, he came up with a better design by joining several cells together. A *battery* is a group of two or more electric cells connected to each other in series, like railway cars in a train. Volta's first battery consisted of several bowls of brine (aqueous sodium chloride) connected by metals that dipped from one bowl into the next (**Figure 1**). This arrangement of metal strips and electrolytes produced a steady flow of electric current.

Figure 1
A version of Volta's first battery. Each beaker contains two different metals, copper and zinc, in an electrolyte, salt water. A series of beakers forms a series of cells (a battery) whose total voltage is the sum of the individual voltages of all cells.

Volta improved the design of this battery by replacing the strips of metal with flat sheets, and replacing the bowls with paper or leather soaked in brine. This produced more electric current for a longer time. As shown in **Figure 2**, Volta stacked cells on top of each other to form a battery, known as a voltaic pile. When a loop of wire was attached to the top and bottom of this voltaic pile, a steady electric current flowed. Volta assembled voltaic piles containing more than 100 cells.

Volta's invention was an immediate technological success because it produced an electric current more simply and more reliably than methods that depended on static charges. It also produced a steady electric current—something no other device could do. The development of this technology led to many advances in physics (for example,

DID YOU KNOW?

Shocking Personal Experiments

"I introduced into my ears two metal rods with rounded ends and joined them to the terminals of the apparatus. At the moment the circuit was completed, I received a shock in the head—and began to hear a noise—a crackling and boiling. This disagreeable sensation, which I feared might be dangerous, has deterred me so that I have not repeated the experiment."

Alessandro Volta (1745–1827)

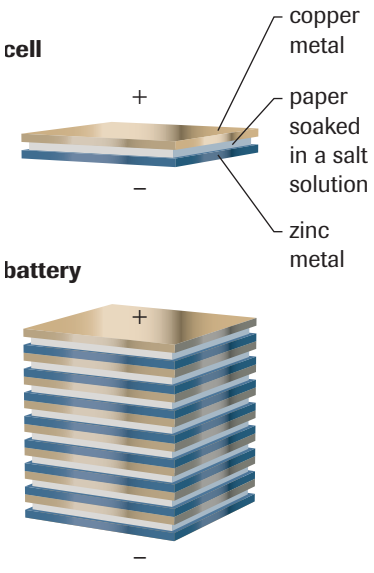

Figure 2
Volta's revised cell design, simpler than the first, consisted of a sandwich of two metals separated by paper soaked in salt water (the electrolyte). A cell consisted of a layer of zinc metal separated from a layer of copper metal by the brine-soaked paper. A large pile of cells could be constructed to give more electrical energy per unit charge.

the theory and description of current electricity), in chemistry (for example, the discovery of Groups 1 and 2 metals), and in electrical and chemical engineering.

Basic Cell Design and Properties

Each electric cell is composed of two **electrodes** (solid electrical conductors) and one **electrolyte** (aqueous electrical conductor) (**Figure 3**). In the cells we buy for home use, the electrolyte is usually a moist paste, containing only enough conducting solution to make the cell function. The electrodes are usually two metals, or graphite and a metal. In some designs, one of the electrodes is the container of the cell. One of the electrodes is marked positive (+) and the other is marked negative (−). *In an electric cell or battery, the cathode is the positive electrode and the anode is the negative electrode.*

According to theory, electricity is the flow of electrons. Electrons move from the anode of a battery through the external circuit to the cathode. A battery produces electricity only when there is an external conducting path, such as a wire, through which electrons can move. Disconnecting the wire from the battery immediately stops the flow of electrons.

A voltmeter is a device that can be used to measure the energy difference, per unit electric charge, between any two points in an electric circuit (**Figure 4(a)**). The energy difference per unit charge is called the **electric potential difference** or the voltage, and is measured in volts (V). For example, the electrons transferred via a 1.5 V cell release only one-sixth as much energy compared with the electrons from a 9 V battery.

Since the voltage is a ratio of energy to charge, it is not dependent on the size of the cell. You may have noticed that you can buy the same type and brand of 1.5 V cells in a variety of sizes, such as AA, B, C, and D. All are rated at 1.5 V. The larger cells can store more energy at the same time as transferring more charge, but the ratio of energy to charge is the same as the smaller cells. *The voltage of a cell depends mainly on the chemical composition of the reactants in the cell.*

Electric current, measured by an ammeter in amperes (A), is a measure of the rate of flow of charge past a point in an electrical circuit (**Figure 4(b)**). The larger the electric cell of a particular type, the greater the current that can be produced by the cell. The charge transferred by a cell or battery is measured in coulombs (C) and expresses the total charge transferred by the movement of charged particles.

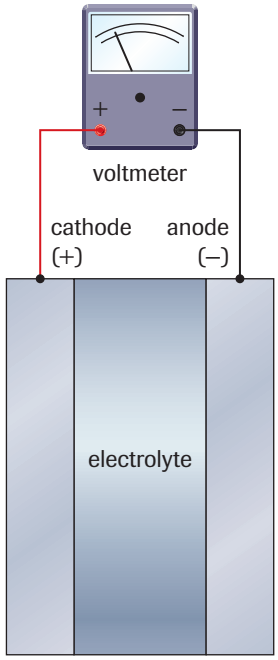

Figure 3
An electric cell always contains two electrodes—an anode and a cathode—and an electrolyte. When testing the voltage of a cell or battery, the red (+) lead of the voltmeter is connected to the positive electrode (cathode), and the black (−) lead is connected to the negative electrode (anode).

(a) **(b)**

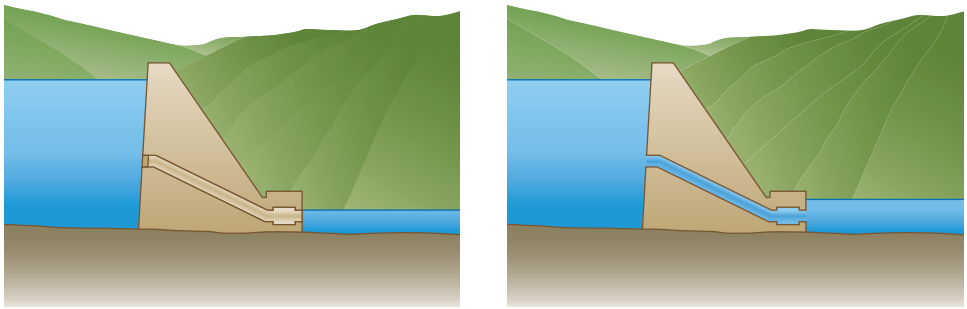

Figure 4
(a) A dam built across a stream or river stops the flow of water. There is a potential energy difference between a kilogram of water at the top of the dam and a kilogram of water at the bottom of the dam. A disconnected cell or battery is like a dam holding back electrons. There is a potential energy difference between the electrons at the anode and the electrons at the cathode. A voltmeter measures this potential energy difference.
(b) If water is released from behind the dam, it naturally flows from the region of higher potential energy (behind the dam) to a lower potential energy below the dam. Similarly, when an electric circuit is connected with a cell or battery, the electrons naturally flow because there is a difference in potential energy between the anode and the cathode.

DID YOU KNOW ?

Electric Charge and Current Analogy

We can extend the analogy in **Figure 4**. The mass (in kg) of water stored behind the dam is similar to the available charge (in C) stored in the chemicals of a cell. When released, the current, or flow, of water (in kg/s) is similar to the current, or flow, of electrons (in C/s).

The *power* of a cell or battery is the rate at which it produces electrical energy. Power is measured in watts (W), and is calculated as the product of the current and the voltage of the battery. The *energy density*, or *specific energy* of a battery, is a measure of the quantity of energy stored or supplied per unit mass. Energy density may be measured in joules per kilogram (J/kg). **Table 1** summarizes some important electrical quantities and their units of measurement.

Table 1 Electrical Quantities and SI Units

Quantity	Symbol	Meter	Unit	Unit symbol
charge	Q	–	coulomb	C
current	I	ammeter	ampere	A (1 A = 1 C/s)
potential difference	V	voltmeter	volt	V (1 V = 1 J/C)
power	P	–	watt	W (1 W = 1 J/s)
energy density	–	–	joules per kilogram	J/kg

SUMMARY Components of an Electric Cell

- An electric cell must have two electrodes and an electrolyte.
- An electrode is a solid conductor.
- An electrolyte is an aqueous conductor.
- The cathode is the electrode labelled positive.
- The anode is the electrode labelled negative.
- The electrons flow through the external circuit from the anode to the cathode.

Practice

1. What are the components of a simple electric cell?
2. Write an empirical definition of electrode and electrolyte, and a conventional definition of anode and cathode.
3. If a DVD player requires 9 V to operate, how many 1.5 V "dry" cells connected in series would it need?
4. Differentiate between electric current and voltage.
5. Why do manufacturers of battery-operated devices print a diagram showing the correct orientation of the batteries? (Supply two answers to this question: one from a scientific perspective and one from a technological perspective.)
6. List several examples illustrating how a new technology (electric cells) led to new scientific discoveries.

DID YOU KNOW ?

Aboriginal Science and Technology

Aboriginal peoples have a long history of technological development in many areas such as agriculture, food preservation, medicine, and transportation. Just as in Western societies, Aboriginal technology had a practical purpose, was developed through a trial-and-error process, and was interdependent on Aboriginal science using traditional or indigenous knowledge.

Technological Problem Solving

The initial development of cells and batteries preceded much of the current scientific understanding of these devices. Cells and batteries existed almost 100 years before the electron was discovered. The study of electric cells is a good illustration of tremendous advances in technology based on very limited scientific knowledge. Technological development or problem solving is similar in some ways to scientific problem solving, but

its purpose differs. The purpose of technological problem solving is to find a realistic way around a practical difficulty—to make something work—while the purpose of scientific problem solving is to describe, explain, and ultimately understand natural and technological phenomena. Technology and science are dependent on each other. Although scientific knowledge can be used to guide the creation of a technology, the technology may create new scientific understanding.

A systematic trial-and-error process is often used in technological problem solving (Appendix C.1). This is not a new process. Aboriginal peoples used a systematic trial-and-error process to ensure the survival of the tribe. An example of a trial-and-error process is as follows:

- Develop a general design for problem-solving trials; for example, select which variables to manipulate and which to control.
- Follow several prediction–procedure–evidence–analysis cycles, manipulating and systematically studying one variable at a time.
- Complete an evaluation based on criteria such as efficiency, reliability, cost, and simplicity.

This technological problem-solving model was important in the early development of practical electric cells.

DID YOU KNOW ?

A "Not Quite Dry" Cell

The electrolyte in the "dry cell" is actually a moist paste (**Figure 5**). If the cell were completely dry, it would not work because the ions in the electrolyte must be able to carry the electric current to complete the circuit. Just enough water is added so that the ions can move, but not enough to make the mixture liquid.

Figure 5

INVESTIGATION 14.1 *Introduction*

Designing an Electric Cell

In this cell, an aluminium soft-drink can is one of the electrodes. The other electrode is a solid conductor, such as a piece of copper wire or pipe, an iron nail, or graphite from a pencil. The electrolyte may be a salt solution, or an acidic or basic solution. Although many characteristics of a cell are important, only one characteristic, voltage, is investigated here. Check with your teacher if you want to evaluate other designs and materials.

When evaluating the Purpose (Part 3), include your opinion about the reliability, cost, and simplicity of your final electric cell.

Purpose

The purpose of this investigation is to make an electric cell.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	● Materials	● Evaluation (1, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Problem

What combination of electrodes and electrolyte gives the largest voltage for an aluminium-can cell?

Design

In a trial-and-error procedure, different variables (second electrode, electrolyte) are modified one at a time while all other variables are held constant. The voltage of each cell is measured as the responding variable.

To perform this investigation, turn to page 658.

Consumer, Commercial, and Industrial Cells

Since Volta's invention of the electric cell and battery, there have been many advances in electrochemistry and technology. Invented in 1865, the zinc chloride cell is commonly referred to as a dry cell because this design was the first to use a sealed container. The first 9 V battery was made up of a series of 1.5 V dry cells (**Figure 6**). Both the 1.5 V dry cell and the 9 V battery are simple, reliable, and relatively inexpensive. Other cells, such as the alkaline dry cell and the mercury cell (**Table 2** on page 616), were developed to improve the performance of the original dry cell. One problem with all of these cells is that the chemicals eventually become depleted and irreversible reactions prevent these cells from being recharged. Cells that cannot be recharged are called *primary cells*.

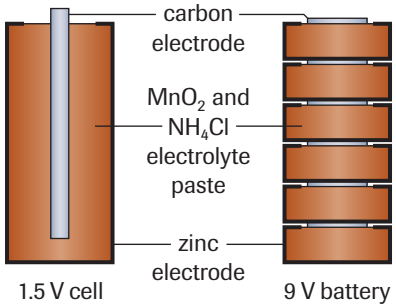

Figure 6

Like a flashlight D cell, the zinc chloride dry cell on the left has a voltage of 1.5 V. The 9 V battery on the right is made up of six 1.5 V dry cells in series.

Table 2 Primary, Secondary, and Fuel Cells

Type	Name of cell	Half-reactions	Characteristics and uses
primary cells	dry cell (1.5 V)	$2\,MnO_2(s) + 2\,NH_4^+(aq) + 2\,e^- \rightarrow Mn_2O_3(s) + 2\,NH_3(aq) + H_2O(l)$ $Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$	• inexpensive, portable, many sizes • flashlights, radios, many other consumer items
	alkaline dry cell (1.5 V)	$2\,MnO_2(s) + H_2O(l) + 2\,e^- \rightarrow Mn_2O_3(s) + 2\,OH^-(aq)$ $Zn(s) + 2\,OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2\,e^-$	• longer shelf life; higher currents for longer periods compared with dry cell • same uses as dry cell
	mercury cell (1.35 V)	$HgO(s) + H_2O(l) + 2\,e^- \rightarrow Hg(l) + 2\,OH^-(aq)$ $Zn(s) + 2\,OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2\,e^-$	• small cell; constant voltage during its active life • hearing aids, watches
secondary cells	Ni–Cd cell (1.25 V)	$2\,NiO(OH)(s) + 2\,H_2O(l) + 2\,e^- \rightarrow 2\,Ni(OH)_2(s) + 2\,OH^-(aq)$ $Cd(s) + 2\,OH^-(aq) \rightarrow Cd(OH)_2(s) + 2\,e^-$	• can be completely sealed; lightweight but expensive • all normal dry cell uses, as well as power tools, shavers, portable computers
	lead–acid cell (2.0 V)	$PbO_2(s) + 4\,H^+(aq) + SO_4^{2-}(aq) + 2\,e^- \rightarrow PbSO_4(s) + 2\,H_2O(l)$ $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2\,e^-$	• very large currents; reliable for many recharges • all vehicles
fuel cells	hydrogen–oxygen cell (1.2 V)	$O_2(g) + 2\,H_2O(l) + 4\,e^- \rightarrow 4\,OH^-(aq)$ $2\,H_2(g) + 4\,OH^-(aq) \rightarrow 4\,H_2O(l) + 4\,e^-$	• lightweight; high efficiency; can be adapted to use hydrogen-rich fuels • vehicles and space shuttle
	aluminium–air cell (2 V)	$3\,O_2(g) + 6\,H_2O(l) + 12\,e^- \rightarrow 12\,OH^-(aq)$ $4\,Al(s) \rightarrow 4\,Al^{3+}(aq) + 12\,e^-$	• very high energy density; made from readily available aluminium alloys • designed for electric cars

Secondary Cells

Secondary cells can be recharged by using electricity to reverse the chemical reaction that occurs when electricity is produced by the cell. Secondary cells and batteries include the nickel–cadmium (Ni–Cd) cell and the lead–acid battery (**Table 2** and **Figure 7**). A relatively recently developed secondary cell with a unique design is a lithium-ion cell, called the Molicel.

EXTENSION

The Molicel

Find out about a made-in-Canada technological development that has had a big impact on portable power.

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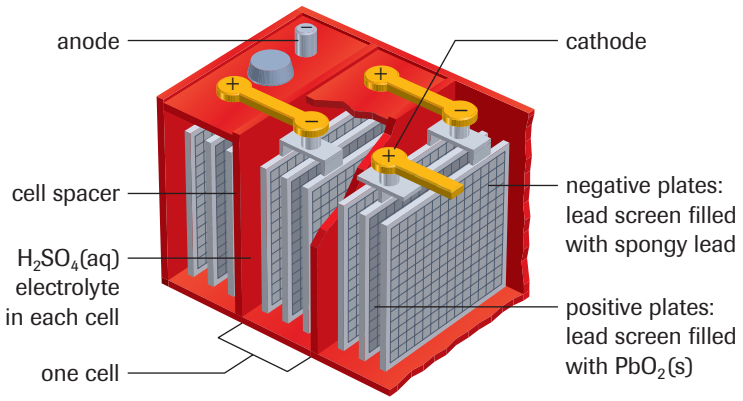

Figure 7
The anodes of a lead–acid car battery are composed of spongy lead and the cathodes are composed of lead(IV) oxide on a metal screen. The large electrode surface area is designed to deliver sufficient current to start a car engine.

One of the most common and reliable secondary cells is the lead–acid cell in a typical car battery. The discharging of this cell (see lead–acid cell, **Table 2**) produces approximately 2.0 V based on the following net equation.

$$Pb(s) + PbO_2(s) + 2\,H_2SO_4(aq) \xrightarrow{\text{discharging}} 2\,PbSO_4(s) + 2\,H_2O(l)$$

This cell requires the input (from the car's alternator) of at least 2.0 V to force the products to change back to the reactants to recharge it. The half-reactions for the lead–acid cell listed in **Table 2** must both be reversed to obtain the following net equation.

$$2\,PbSO_4(s) + 2\,H_2O(l) \xrightarrow{\text{charging}} Pb(s) + PbO_2(s) + 2\,H_2SO_4(aq)$$

A battery can be recharged if the products are stable with no further reactions occurring and if the products are able to travel through the electrolyte toward the appropriate electrode.

Learning Tip

Discharging a cell or battery is like letting the water spontaneously run out from the higher level behind a dam. Charging (or recharging) is like pumping the water up behind the dam. This is not a spontaneous process and requires energy.

Fuel Cells

A fuel cell is a different solution to the problem of the limited life of a primary cell. **Fuel cells** produce electricity by the reaction of a fuel that is continuously supplied to keep the cell operating. In principle, the fuel cell could be used forever, provided the fuel is continuously supplied. The fuel cell offers several advantages over methods that produce electricity by the combustion of fossil fuels. For example, fuel cells generate electricity more efficiently (**Table 3**), without producing greenhouse gases or substances that contribute to acid rain. The development of a cost-effective fuel cell is currently the focus of much scientific study and technological research and development.

William Grove accidentally invented the first fuel cell in 1839 using platinum electrodes, hydrogen, and oxygen as fuels, and sulfuric acid as the electrolyte. Grove was actually studying the reverse process—using electricity to convert water into hydrogen and oxygen. After one experiment, he reconnected the two electrodes without a power supply attached, and found that a small current was produced spontaneously as hydrogen and oxygen combined to form water. Grove continued to work on this cell, but eventually decided it was not a practical device because the characteristics, such as voltage, current, and energy capacity, were quite poor. Although many scientists, including Nobel Prize winners Fritz Haber and Walther Nernst, worked on improving the cell, they were largely unsuccessful. They manipulated many variables, such as different electrodes and electrolytes, but the reaction rates were too low and the electrodes became corroded.

Finally, in 1955, Francis Bacon succeeded where many others had failed. He produced a practical hydrogen–oxygen fuel cell using an alkaline electrolyte and electrodes constructed of porous nickel (**Figure 8**). Although the idea had been around for a long time, Bacon's cell was really the first practical fuel cell. NASA quickly adopted the hydrogen–oxygen fuel cell as an electrical power source for space flights because hydrogen and oxygen are already available for propulsion systems and the product, water, can be purified for drinking. NASA's fuel cell, a modification of the original Bacon cell, is an alkaline cell using potassium hydroxide as the electrolyte (**Table 2** on page 616). It produces 12 kW of electricity and operates at 70% efficiency. Unfortunately, NASA's fuel cell is expensive and has a relatively short working life, primarily due to the corrosive electrolyte. As a result, NASA's cell is not economically viable for general or commercial applications.

Table 3 Efficiencies of Different Technologies*

Technology	Efficiency*
fuel cells	40–70%
electric power plants	30–40%
automobile engines	17–23%
gasoline lawn mower	about 12%

*Efficiency is the fraction of the maximum available energy that is actually usable.

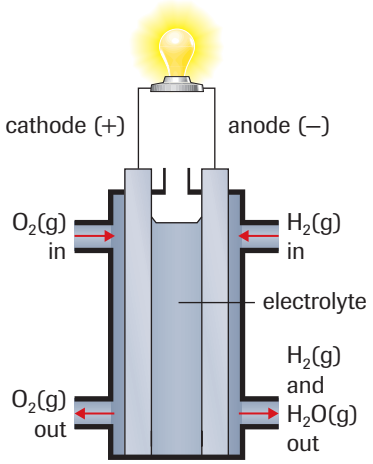

Figure 8
Hydrogen and oxygen gases are continuously pumped into this hydrogen–oxygen fuel cell. Each gas reacts at a different electrode. Unused gases are recycled.

EXTENSION

CBC radioONE

QUIRKS & QUARKS

Fuel Cells

Car designers are working on futuristic cars, powered by fuel cells, that would cut down on pollution. This video shows practical applications of science and technology that may, some day, bring us cheap, clean, efficient transportation.

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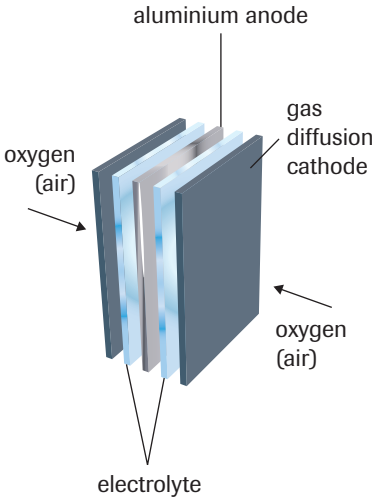

Figure 9
Several Canadian companies have done extensive research in the development of the aluminium–air solid fuel cell.

WEB Activity

Web Quest—Hydrogen: Wonderfuel or Hype?

Taking the role of a research company, you and your group will investigate and prepare an illustrated presentation on the subject of hydrogen as a fuel. Is it clean? Is it inexpensive? Is it widely available? Will hydrogen become the wonderfuel of the 21st century or is hydrogen just power hype?

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Aluminium–Air Cell

Another type of fuel cell is the metal–air fuel cell, the most common of which is the aluminium–air cell (**Table 2**). This is actually an aluminium–oxygen cell and has been developed for possible use in electric cars. Air is pumped into the cell and oxygen reacts at the cathode while a replaceable mass of aluminium reacts at the anode (**Figure 9**). The fuel is solid aluminium metal and the product, aluminium hydroxide, can be recycled back to aluminium metal. The simple design means that this cell can be assembled in almost any size. The high energy density of these cells results from the fact that three moles of electrons are released from every mole of aluminium, and aluminium is a very lightweight metal. Unlike hydrogen, storage and transportation of a solid fuel do not pose a problem. Estimates from prototypes suggest that the aluminium anode will need replacement every 2500 km in an electric car.

Large-Scale Commercial and Industrial Fuel Cells

The requirements for electrical power fuel cells for large-scale use in businesses and industry are similar with regard to the fuel, but there is less concern about volume, weight, or energy density. However, businesses and industries need cells with much longer lifetimes. Fuel cells for large-scale commercial and industrial use are almost always co-generation units. This means that they produce electricity as well as heat for space heating. Co-generation means that the overall efficiencies can be as high as 90%. Commercially viable fuel cells today are usually acid electrolyte cells such as the phosphoric acid fuel cell, which can produce 400 MW of power, sufficient for meeting the electrical energy needs of a small city (**Figure 10**). These cells usually use natural gas as a source of hydrogen for the fuel cell and operate at temperatures of 200 °C.

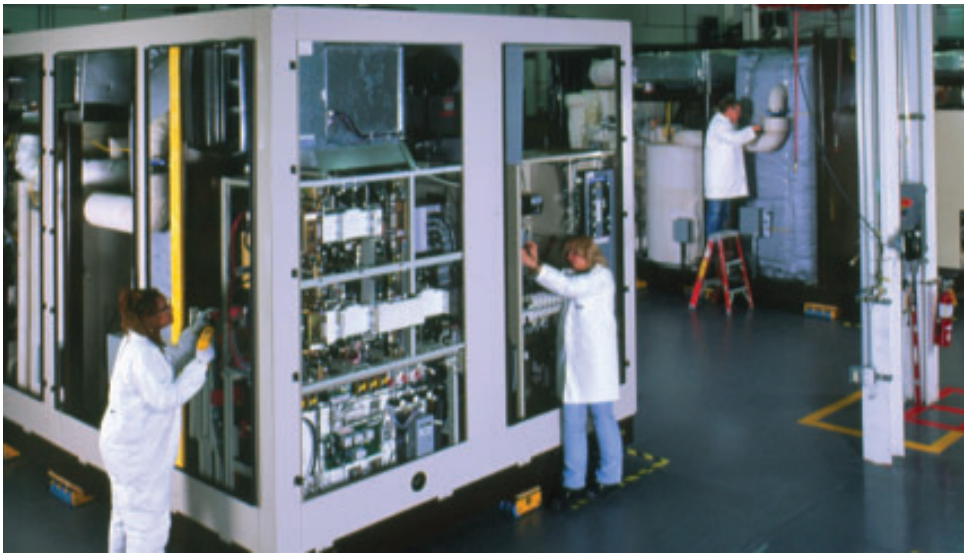

Figure 10
The world's first commercial phosphoric acid fuel cell was produced by ONSI/International Fuel Cells. It has been available since 1992 and uses natural gas, waste methane, propane, or hydrogen as fuels. This unit produces 200 kW electricity and 200 kW heat at a total system efficiency of 80%.

Practice

7. Compare scientific and technological problem solving.
8. What steps are involved in technological problem solving?
9. Suppose you decided to develop and market an aluminium-can cell (see Investigation 14.1.) How and why would you alter the electrolyte?
10. Distinguish between primary and secondary cells, including a common example of each.
11. What are some advantages and disadvantages of the zinc chloride dry cell?
12. What do the designs of a dry-cell container and an ice-cream cone have in common?
13. One of the most successful batteries has been the lead–acid car battery.
    (a) Identify the anode, cathode, and electrolyte.
    (b) How are the large currents produced that are necessary to start a car?
    (c) What has been the social impact of this battery?
    (d) What are some possible environmental impacts of this battery?
14. For both the hydrogen–oxygen fuel cell and the aluminium–air fuel cell,
    (a) write the two half-reaction equations (refer to **Table 2**, page 616)
    (b) label each equation from (a) as an oxidation or a reduction
    (c) write the net ionic equation for each cell
15. Using several perspectives, state some advantages and disadvantages of a fuel cell.
16. Assess the possible future importance of fuel cells in society.
17. Experimental cell phones that run on miniature hydrogen–oxygen fuel cells exist today. What must be in place before the average consumer can buy one?

DID YOU KNOW?

Success or Failure?

Thomas Edison, the American inventor, set out to invent a secondary battery. His plan was to use an alkaline electrolyte and iron as the anode, but he needed to find a suitable cathode material. After thousands of experiments, his friend and associate Walter Mallory commented on Edison's lack of results. Edison replied that he had lots of results and he knew fifty thousand things that did not work! Eventually he discovered that a thin metallic nickel film produced a battery that was lighter and more powerful than the existing lead–acid batteries.

WEB Activity

Canadian Achievers—Lewis Urry

After graduating in chemical engineering from the University of Toronto in 1950, Lewis Urry (**Figure 11**) went to work for Eveready Battery Company where he developed the first practical, long-life electric cell.

1. Identify the unique feature in Urry's cell ("battery") and explain how it works.
2. List three familiar devices that depend on the alkaline cell.
3. Name one other cell developed by Urry.

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Figure 11
Lewis Urry (1927–2004), inventor of the alkaline battery

The Ballard Fuel Cell

A variation of a hydrogen–oxygen fuel cell, also known as the hydrogen fuel cell, was developed for commercial applications by Ballard Power Systems in Burnaby, BC. The Ballard fuel cell uses a proton exchange membrane (PEM) in place of a liquid electrolyte. Normal electric cells use the ions in the liquid electrolyte to transfer electric charge within the cell. In a hydrogen fuel cell, the PEM is made from a solid proton-conducting polymer that transfers charge within the cell (**Figure 12**). The PEM is simple, robust, eliminates corrosive liquids, and permits a high energy density.

The Ballard fuel cell consists of an anode and a cathode separated by a PEM. Hydrogen fuel admitted through a porous anode is converted into hydrogen ions (protons) and free electrons in the presence of a catalyst at the anode. An external circuit conducts the free electrons and produces the desired electrical current. Water and heat are produced when the protons, after migrating through the polymer membrane to the cathode, react both with oxygen molecules from the air and with the free electrons from the external circuit.

Fuel cells can be connected in series (stacked) to increase the voltage and power output. For example, an experimental transit bus uses an electric motor powered by a Ballard fuel cell that is capable of 205 kW (or 250 hp).

Ballard has development agreements with most major car manufacturers to use its cells in future electric cars. The zero-emission engines convert hydrogen, or hydrogen-rich fuels such as natural gas and even methanol, into electricity, producing water and heat as the main byproducts.

Although the Ballard hydrogen fuel cell looks very promising, there are several problems yet to be solved. Cost is a major factor, which may be partially solved by mass production. The fuel is also under debate. If hydrogen gas is used, where does it come from? Electrolysis of water uses a lot of electrical energy and, if this energy comes from a fossil-fuel generating plant, a lot of pollution is produced along with the hydrogen gas. How would the hydrogen gas be distributed and stored on board the vehicle? There are important safety concerns associated with the handling and storage of hydrogen, which is flammable. Many scientists and engineers believe that the solution is not to use hydrogen gas directly, but to use hydrogen-rich fuels. We have a lot of knowledge of reforming hydrocarbons to produce hydrogen. If natural gas or even gasoline were reformed as needed on board the vehicle, then we would have a familiar fuel source and an infrastructure in place to supply this fuel. Not everyone agrees that this is a good solution.

Hydrogen fuel cells may change the way that we use energy, especially for transportation. Fuel cells may be used not only to power cars, buses, and boats, but also many portable devices.

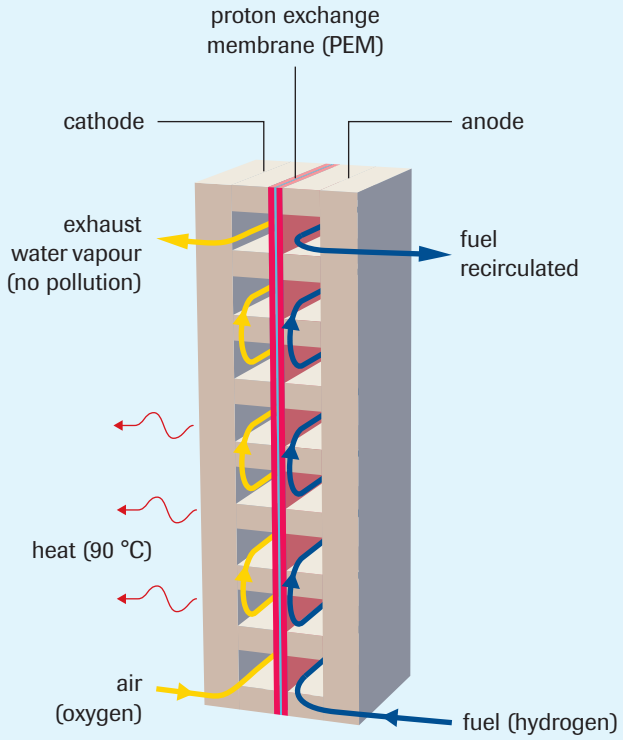

Figure 12
The hydrogen fuel cell has the same design as Volta's original cell, but the electrolyte is a conducting solid.

Case Study Questions

1. Is hydrogen a "clean" fuel? Explain your answer.
2. Make a list, in order of importance, of the advantages of a Ballard hydrogen fuel cell for urban buses compared with those of a diesel engine for buses. Make a similar list of possible disadvantages.
3. How has the development of fuel cells been influenced by society and how might they influence society in the future? Include a variety of perspectives. Use information from various print and electronic sources.

Extension

4. One solution to the hydrogen storage problem is to make hydrogen as it is needed using a reformer. Research how reformers can be used to produce hydrogen.

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5. Iceland has ambitious plans to be the first nation to replace all its fossil fuels with "clean" hydrogen. Research how Iceland plans to produce the hydrogen it will need.

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Section 14.1 Questions

1. Draw a simple diagram of an electric cell and label: electrodes, electrolyte, cathode, anode, signs for cathode and anode, and direction of electron flow through an external wire.
2. State the evidence that an electric cell involves a redox reaction.
3. List three types of electric cells used in consumer and commercial operations. Briefly describe the main feature of each cell.
4. State two common examples of consumer cells and where they may be used.
5. A silver oxide cell is often used when a miniature cell or battery is required, as in watches, calculators, and cameras. The following half-reaction equations occur in the cell:

   $Ag_2O(s) + H_2O(l) + 2\,e^- \rightarrow 2\,Ag(s) + 2\,OH^-(aq)$

   $Zn(s) + 2\,OH^-(aq) \rightarrow Zn(OH)_2(s) + 2\,e^-$

   (a) In which direction does the electric current flow: silver to zinc or zinc to silver?
   (b) Which is the anode and which is the cathode?
   (c) Write the net redox equation for the discharging of the silver oxide cell.
6. Describe an example where a new technology led to scientific discoveries and an example in which scientific knowledge led to a better technology.
7. Possible solutions to technological problems often involve different designs, materials, and processes. List several examples that solve the problem of cells becoming depleted and unusable.
8. What are some unintended consequences of the widespread use of many different types of cells and batteries?
9. Describe how societal needs and expectations have affected the development of the electric cells available today.
10. Suppose cells and batteries did not exist. What impact would that have on your life?
11. State several reasons why it is important to recycle the components of cells and batteries when they have reached the end of their useful life.
12. (a) Why is there a great deal of interest in electric cars?
    (b) Suggest some reasons why we don't use lead–acid batteries as the only power source for electric cars.
    (c) How have advances in hydrogen fuel cells facilitated the development of electric cars?

Extension

13. Most people associate technological development with "progress" in Western societies and think that technology transfer to Aboriginal peoples only occurs in one direction. In fact, although not often recognized and valued, there has been significant technology transfer from Aboriginal peoples to Western society. Summarize several examples of significant Aboriginal technological knowledge that was transferred to early European settlers. What area is still a significant source of technological transfer from Aboriginal to Western societies? Why is this an area of dispute?

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14. Portable electronic devices can be found everywhere. Laptop computers, cellular telephones, mobile radios, cordless phones, portable disc and MP3 players, and digital cameras all require an electric cell.
    (a) What are some of the requirements for cells used in these applications?
    (b) Why are some rechargeable batteries used in various portable devices supposed to be totally "drained" (discharged) before recharging?

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15. People whose heart occasionally beats too slowly or too quickly often have a pacemaker fitted to keep the heart beating regularly (**Figure 13**). Pacemakers use a battery for electric power. What kind of battery is commonly used today? How long does it last? How does the doctor know when the battery is nearing the end of its life and needs to be replaced? Why are rechargeable batteries generally not used?

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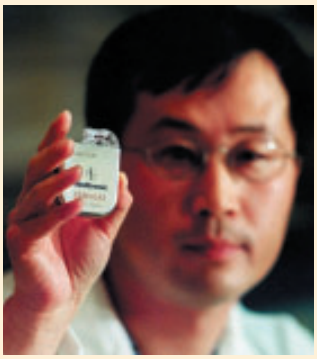

Figure 13
A pacemaker contains electronics and a built-in battery. The whole unit is only a few centimetres in size and is implanted under the skin near the collarbone.

16. Plastic batteries were the dream of the 1980s, the disappointment of the 1990s, and the subject of the 2000 Nobel Prize for Chemistry. Now it appears that some commercial products will eventually result from the research and development invested in plastic batteries. Briefly describe the electrodes and electrolyte for a plastic battery. How is this battery similar to and different from an ordinary battery? What are some advantages and disadvantages?

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14.2 Voltaic Cells

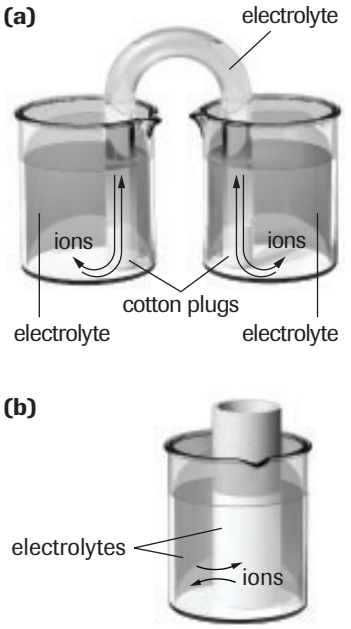


Figure 1
(a) A salt bridge is a U-shaped tube containing an inert (unreactive) aqueous electrolyte such as sodium sulfate.
(b) An unglazed porcelain (porous) cup containing one electrolyte sits in a container of a second electrolyte.

Electric cells were invented for practical purposes in about 1800, but they were not explained scientifically until about 100 years after their invention. Their use, however, contributed to the scientific understanding of redox reactions. Later this knowledge helped explain reactions inside the cell itself. Electric cells adapted for scientific study are often called *galvanic cells* (in recognition of Luigi Galvani), or *voltaic cells* (in recognition of Alessandro Volta).

You learned in Chapter 13 that a redox reaction involves a transfer of electrons from the reducing agent to the oxidizing agent. In a voltaic cell, electrons pass from the reducing agent to the oxidizing agent through an external circuit rather than passing directly from one substance to another. You have seen that the individual components of a cell—electrodes and electrolytes—determine characteristics such as voltage and current. Why is this so? What happens in different parts of a cell? To answer these questions, chemists use a cell that has two electrodes and their electrolytes separated. This is not a very practical arrangement, but it greatly facilitates the study of cells. Each electrode is in contact with an electrolyte. A **porous boundary** separates the two electrolytes, at least for a short time, while still permitting ions to move between the two solutions through tiny openings in the cotton plugs of a salt bridge (**Figure 1(a)**) or in the walls of a porcelain cup (**Figure 1(b)**).

Using this design modification, a cell can be split into two parts connected by a porous boundary. Each part, called a **half-cell**, consists of one electrode and one electrolyte. For example, the copper–zinc cell shown in **Figure 2** has two half-cells, copper metal in a solution of copper ions and zinc metal in a solution of zinc ions.

This cell can be represented using the following abbreviated ("shorthand") notation, called a cell notation:

$$Zn(s) \mid Zn(NO_3)_2(aq) \parallel Cu(NO_3)_2(aq) \mid Cu(s)$$

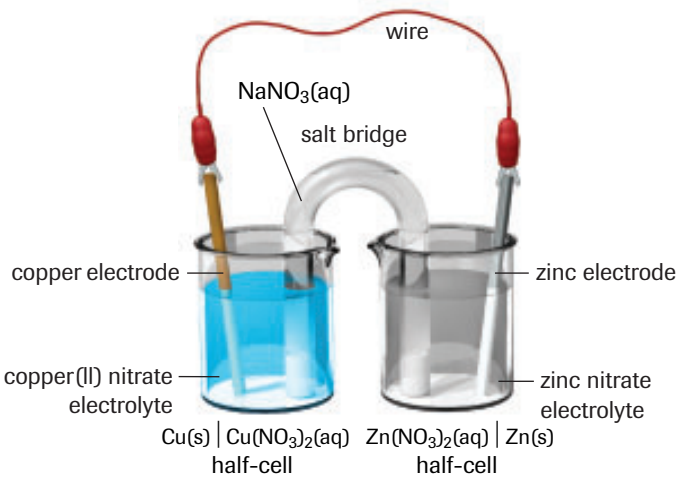


Figure 2
Each electrode is in its own electrolyte, forming a half-cell. The two half-cells are connected by a salt bridge (containing $NaNO_3(aq)$) and by an external wire to make a complete circuit.

In this notation, a single line (|) indicates a phase boundary, such as the interface of an electrode and an electrolyte in a half-cell. A double line (||) represents a physical boundary, such as a porous boundary between half-cells. A **voltaic cell** is an arrangement of two half-cells separated by a porous boundary. Voltaic cells, such as the one in **Figure 2**, are especially suitable for scientific study.

Learning Tip

It is a common practice to write the cell notation with the anode on the left-hand side and the cathode on the right-hand side. Diagrams or photos of cells can have any arrangement.

INVESTIGATION 14.2 ***Introduction***

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ● Evidence
- ● Analysis
- ○ Evaluation

A Voltaic Cell (Demonstration)

An important characteristic of consumer, commercial, and industrial cells is a simple, efficient design that works for the intended application. For scientific research, this is not as important as a design that can be easily manipulated and studied.

Purpose

The purpose of this investigation is to test the design and operation of a voltaic cell used in scientific research.

Problem

What is the design and operation of a voltaic cell?

Design

An electric cell with only one electrolyte is compared with voltaic cells containing the same electrodes, but two electrolytes.

To perform this investigation, turn to page 659.

A Theoretical Description of a Voltaic Cell

Observation of a voltaic cell as it operates provides evidence to explain what is happening inside the cell. For example, during your study of the silver–copper cell in Investigation 14.2, you observed the evidence listed in **Table 1**. Note that the table also includes a theoretical interpretation of each point, which is also shown in **Figure 3**.

According to the electron transfer theory and the concept of relative strengths of oxidizing and reducing agents, silver ions are the strongest oxidizing agents in the cell; they undergo a reduction half-reaction at the cathode. The strongest oxidizing agent in the cell always undergoes a reduction at the cathode. Copper atoms, which are the strongest reducing agents in the cell, give up electrons in an oxidation half-reaction and enter the

Learning Tip

People often use acronyms or similar devices to help them remember important information. "OIL RIG (oxidation is loss, reduction is gain)" is an example. One way to help you remember important details of a cell is the expression SOAC/GERC, loosely read as "soak a jerk." Translated, this means the Strongest Oxidizing Agent at the Cathode Gains Electrons and is Reduced at the Cathode. Another example is "An ox ate a red cat," which helps to recall **a**node **o**xidation; **r**eduction **c**athode.

Table 1 Evidence and Interpretations of the Silver–Copper Cell

Evidence	Interpretation
The copper electrode decreases in size and the intensity of the blue colour of the electrolyte increases.	Oxidation of copper metal is occurring: $Cu(s) \rightarrow \underset{\text{blue}}{Cu^{2+}(aq)} + 2\,e^-$
The silver electrode increases in size as long, silver-coloured crystals grow.	Reduction of silver ions is occurring: $Ag^+(aq) + e^- \rightarrow Ag(s)$
A blue colour slowly moves up the U-tube from the copper half-cell to the silver half-cell.	Copper(II) ions (cations) move toward the cathode.
An ammeter shows that the electric current flows along a wire between the copper electrode and the silver electrode.	Electrons move through the wire from the copper electrode to the silver electrode.
A voltmeter indicates that the silver electrode cathode (positive) and the copper electrode is the anode (negative).	Electrons have a tendency to leave the copper half-cell and enter the silver half-cell.

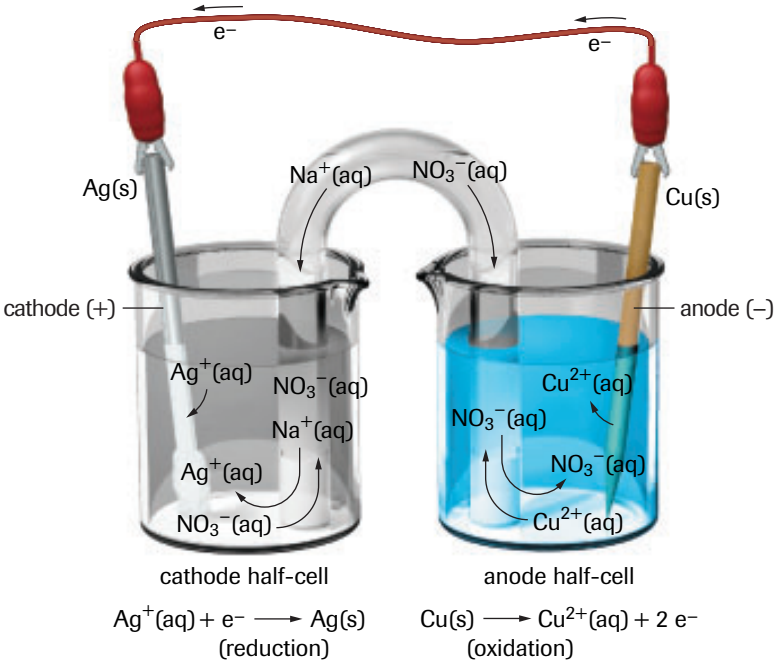


Figure 3
A theoretical interpretation of the silver–copper cell

solution at the anode. The strongest reducing agent in the cell always undergoes an oxidation at the anode. Therefore, the **cathode** is the electrode where reduction occurs and the **anode** is the electrode where oxidation occurs.

- *The strongest oxidizing agent present in the cell always undergoes a reduction at the cathode.*
- *The strongest reducing agent present in the cell always undergoes an oxidation at the anode.*

Electrons released by the oxidation of copper atoms at the anode travel through the connecting wire to the silver cathode. We can explain the direction of electron flow in terms of competition for electrons. According to the Table of Relative Strengths of Oxidizing and Reducing Agents (redox table) in Appendix I, silver ions are stronger oxidizing agents than copper(II) ions. Silver ions win the tug of war for the electrons available from the conducting wire.

To write the net equation for the silver–copper voltaic cell, we first identify the strongest oxidizing and reducing agents present in the mixture. (This is the same procedure you followed when predicting redox reactions in Section 13.2.) We then follow the same procedure for predicting half-reactions in which the two materials are in contact with each other.

$$\underset{\text{SRA}}{\text{Cu(s)}} \;|\; \underset{\text{OA}}{\text{Cu}^{2+}\text{(aq)}} \;||\; \underset{\text{SOA}}{\text{Ag}^{+}\text{(aq)}} \;|\; \underset{\text{RA}}{\text{Ag(s)}}$$

reduction at the cathode: $2\,[\text{Ag}^{+}\text{(aq)} + \text{e}^{-} \rightarrow \text{Ag(s)}]$

oxidation at the anode: $\text{Cu(s)} \rightarrow \text{Cu}^{2+}\text{(aq)} + 2\,\text{e}^{-}$

net: $\text{Cu(s)} + 2\,\text{Ag}^{+}\text{(aq)} \rightarrow \text{Cu}^{2+}\text{(aq)} + 2\,\text{Ag(s)}$

Experimental evidence shows that all parts of the cell remain electrically neutral. You can see from the half-reaction equations that silver ions are removed from the solution when they gain electrons to form silver solid. Because the solution started as electrically neutral, the removal of positively charged ions would mean the solution would become negatively charged. However, based on the evidence, this does not happen. Cations, which in this case are $Na^{+}(aq)$ ions, move from the salt bridge into the solution in the cathode compartment to maintain an electrically neutral solution (refer to **Figure 3**). Similarly, as positively charged copper(II) ions form at the anode, anions, which in this case are $NO_3^{-}(aq)$ ions, move from the salt bridge into the anode compartment to maintain electrical neutrality.

For a general voltaic cell, we can summarize the components and processes for a general voltaic cell using the cell notation and cell diagram shown in **Figure 4**.

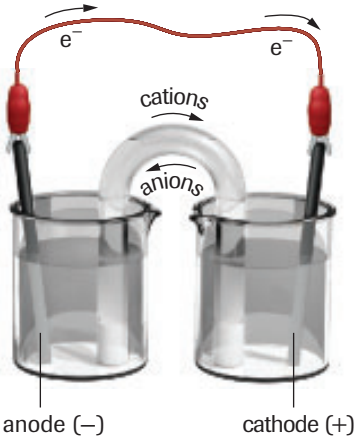
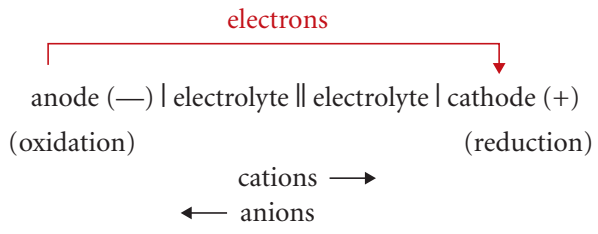


Figure 4
In an operating voltaic cell, the electrical circuit is completed by the electron flow in the external part (wires) of the cell and the ion flow in the internal part (solutions) of the cell.



Voltaic Cells with Inert Electrodes

For cells containing metals and metal ions, the electrodes are usually the metals, and half-reactions take place on the surface of the metals. What happens if an oxidizing or a reducing agent other than these is used? For example, an acidic dichromate solution is a strong oxidizing agent that reacts spontaneously with copper metal. To construct this cell, you can use a copper half-cell, as shown in **Figure 5**, but an electrode is also required for the dichromate half-cell. You cannot use solid sodium dichromate as an electrode because solid ionic compounds do not conduct electricity and solid sodium dichromate would also dissolve in the solution. You need a solid conductor that will not react in the cell or interfere with the desired cell reaction. In other words, you need an unreactive or **inert electrode**. Inert electrodes provide a location to connect a wire and a surface on which a half-reaction can occur. A carbon (graphite) rod or platinum metal foil are two inert electrodes that are commonly used as a cathode or anode.

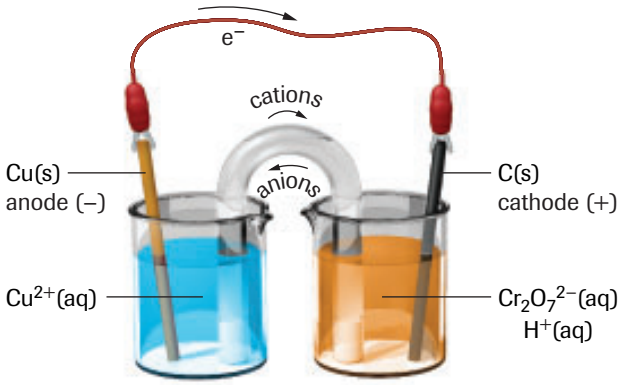


Figure 5
The copper electrode decreases in mass and the blue colour of the electrolyte increases, which indicates oxidation at the anode. The carbon electrode remains unchanged, but the orange colour of the dichromate solution becomes less intense and changes to greenish-yellow, evidence that reduction is occurring in this half cell.

COMMUNICATION example 1

(a) Write equations for the half-reactions and the overall reaction that occur in the following cell:

$$C(s) \mid Fe^{2+}(aq), Fe^{3+}(aq) \parallel Cr_2O_7^{2-}(aq), H^+(aq) \mid C(s)$$

(b) Draw a diagram of the cell, labelling electrodes, electrolytes, the direction of electron flow, and the direction of ion movement.

Solution

(a)

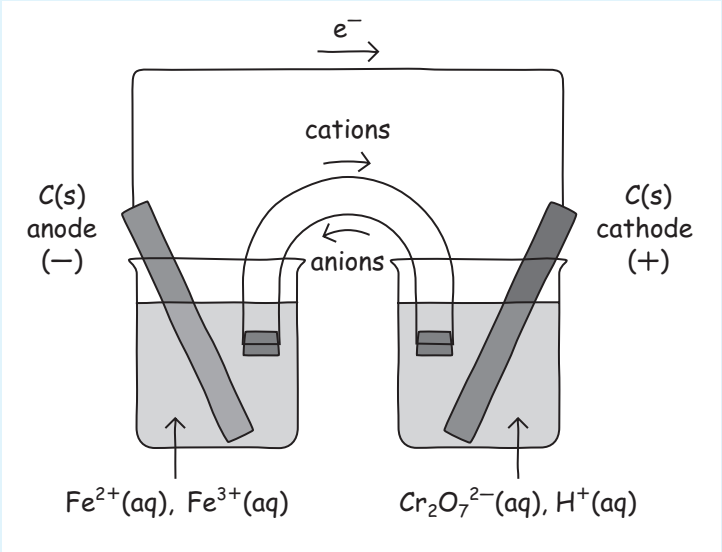
$$\underset{\text{SRA}}{C(s)} \mid \underset{\text{OA}}{Fe^{2+}(aq), Fe^{3+}(aq)} \parallel \underset{\text{SOA}}{Cr_2O_7^{2-}(aq)}, \underset{\text{OA}}{H^+(aq)} \mid C(s)$$

cathode: $Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,e^- \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$

anode: $6\,[\,Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-\,]$

net: $Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,Fe^{2+}(aq) \rightarrow 6\,Fe^{3+}(aq) + 2\,Cr^{3+}(aq) + 7\,H_2O(l)$

(b)



DID YOU KNOW?

Faraday and the Term "Ion"

The term "ion," from the ancient Greek meaning "going" was first used by Michael Faraday in 1834.

> "Finally, I require a term to express those bodies which can pass to the electrodes … I propose to distinguish these bodies by calling those anions which go to the anode of the decomposing body; and those passing to the cathode, cations; and when I have occasion to speak of these together, I shall call them ions."

Learning Tip

A variety of names can be used for cells based upon spontaneous redox reactions: electric, voltaic, galvanic, and electrochemical. In this book, *electric cell* is used for consumer cells and *voltaic cell* is used for scientific research cells. *Electrochemical cell* is a general term that includes both voltaic (spontaneous) and electrolytic (nonspontaneous) cells.

SUMMARY Voltaic Cells

- A voltaic cell consists of two half-cells separated by a porous boundary with solid electrodes connected by an external circuit.
- The cathode is the positive electrode. The strongest oxidizing agent is reduced at the cathode.
- The anode is the negative electrode. The strongest reducing agent is oxidized at the anode.
- Electrons travel in the external circuit from the anode to the cathode.
- Internally, anions move toward the anode and cations move toward the cathode as the cell operates. The solutions remain electrically neutral.
- Cell notation: anode | electrolyte || electrolyte | cathode
  where a single vertical line represents a phase boundary and a double vertical line represents a porous boundary.

DID YOU KNOW?

Electron Sources and Sinks

Chemists sometimes refer to the anode as the electron source and the cathode as the electron sink. Look at some half-reaction equations to see why these terms apply.

Practice

1. Write an empirical description of each of the following terms: voltaic cell, half-cell, porous boundary, salt bridge, electrolyte, external circuit, and inert electrode.
2. Write a theoretical definition of a cathode and an anode.
3. Indicate whether the following processes occur at the cathode or at the anode of a voltaic cell.
   (a) reduction half-reaction
   (b) oxidation half-reaction
   (c) reaction of the strongest reducing agent
   (d) reaction of the strongest oxidizing agent
4. When is an inert electrode used? Give two common examples.
5. What are the characteristics of the solution in a salt bridge? Provide an example.
6. For each of the following cells, use the given cell notation to identify the strongest oxidizing and reducing agents. Write chemical equations to represent the cathode, anode, and net cell reactions. Draw a diagram of each cell, labelling the electrodes, electrolytes, direction of electron flow, and direction of ion movement.
   (a) $Zn(s) \mid Zn^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$
   (b) $Al(s) \mid Al^{3+}(aq) \parallel NO_3^{-}(aq), H^{+}(aq) \mid Pt(s)$
7. Ions move through a porous boundary between the two half-cells of a voltaic cell.
   (a) In what direction do the cations and anions move?
   (b) Why do the ions move? Take your answer and convert it into another "why" question. Now answer this question.
   (c) For the copper–silver cell shown in **Figure 3** (page 624), we know that $Na^{+}(aq)$ and $NO_3^{-}(aq)$ in the salt bridge move into the cathode and anode compartments, respectively. Explain the evidence that the blue colour moves up the salt bridge towards the cathode (refer to **Table 1**, page 623)?
8. Draw and label a diagram for a voltaic cell constructed from some (not all) of the following materials:

strip of cadmium metal	voltmeter
strip of nickel metal	connecting wires
solid cadmium sulfate	glass U-tube
solid nickel(II) sulfate	cotton
solid potassium sulfate	various beakers
distilled water	porous porcelain cup

9. Redesign the voltaic cell in question 8 by changing at least one electrode and one electrolyte. The net reaction should remain the same for the redesigned cell.

Standard Cells and Cell Potentials

The investigations and activities you have completed show that the design and composition of a cell affect its operation. To make comparisons and scientific study easier, chemists specify the composition of a cell and the conditions under which the cell operates. A **standard cell** is a voltaic cell in which each half-cell contains all entities shown in the half-reaction equation at SATP conditions, with a concentration of 1.0 mol/L for the aqueous solutions. If a metal is not part of a half-cell, then an inert electrode is used to construct the standard cell. For example, for a standard zinc–dichromate cell, the cell description is

$$\underset{}{Zn(s)} \mid \underset{1.0\ mol/L}{Zn^{2+}(aq)} \mid\mid \underset{1.0\ mol/L}{Cr_2O_7^{2-}(aq)}, \underset{1.0\ mol/L}{H^+(aq)}, \underset{1.0\ mol/L}{Cr^{3+}(aq)} \mid \underset{}{C(s)} \quad \text{at SATP}$$

The **standard cell potential**, $E°_{cell}$, is the maximum electric potential difference (voltage) of the cell operating under standard conditions; $E°_{cell}$ represents the energy difference (per unit of charge) between the cathode and the anode. The degree sign (°) indicates that standard 1.0 mol/L and SATP conditions apply. Based on the idea of competition for electrons, a **standard reduction potential**, $E°_r$, represents the ability of a standard half-cell to attract electrons, thus undergoing a reduction. The half-cell with the greater attraction for electrons—that is, the one with the more positive reduction potential—gains electrons from the half-cell with the lower reduction potential. The standard cell potential is the difference between the reduction potentials of the two standard half-cells.

$$E°_{cell} = \underset{cathode}{E°_r} - \underset{anode}{E°_r}$$

It is impossible to determine experimentally the reduction potential of a single half-cell because electron transfer requires both an oxidizing agent and a reducing agent. Note that a voltmeter can only measure a potential difference, $E°_{cell}$. To assign values for standard reduction potentials, we measure the "reducing" strength of all possible half-cells relative to an accepted, standard half-cell. The half-cell used for this purpose is the standard hydrogen half-cell. A half-cell such as this, that is chosen as a reference and arbitrarily assigned an electrode potential of exactly zero volts, is called a **reference half-cell**.

Standard Hydrogen Half-Cell

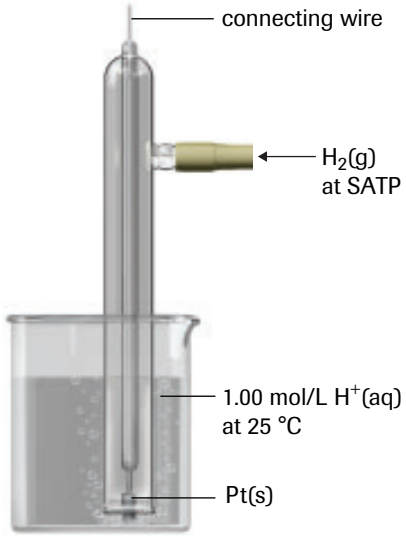
The standard hydrogen half-cell (**Figure 6**) consists of an inert platinum electrode immersed in a 1.00 mol/L solution of hydrogen ions, with hydrogen gas at a pressure of 100 kPa bubbling over the electrode. The pressure and temperature of the cell are kept at SATP conditions. Standard reduction potentials for all other half-cells are measured relative to that of the standard hydrogen half-cell. The reduction potential of the hydrogen ion reduction half-reaction is defined to be exactly zero volts.

$$2\,H^+(aq) + 2\,e^- \rightleftarrows H_2(g) \qquad E°_r = 0.00\ V$$

As a result, we can assign a numerical value to the reduction potential associated with every other reaction. A reduction potential that has a positive value means that the oxidizing agent is a stronger oxidizing agent than hydrogen ions. A negative reduction potential means that the oxidizing agent is a weaker oxidizing agent than hydrogen ions. The choice of the standard hydrogen half-cell as a reference is the accepted convention. If a different half-cell had been chosen as the reference, individual reduction potentials would be different, but their relative values would remain the same.

Learning Tip

Think of the standard cell potential $E°_{cell}$ as representing the *difference* in ability of two half-cells to gain electrons. This potential difference can only be measured accurately if no current is allowed to flow. A good-quality voltmeter has a large internal resistance to prevent current flow.



$Pt(s) \mid H_2(g) \mid H^+(aq) \quad E°_r = 0.00\ V$

Figure 6
The standard hydrogen half-cell is used internationally as the reference half-cell in electrochemical research. Notice that the second vertical line in the cell notation designates the phase boundary between the gas and the liquid.

Learning Tip

A voltmeter has two terminals, positive (red) and negative (black). Connect these to the electrodes of any cell so that the voltmeter gives a positive reading. Whatever electrode is connected to the positive terminal will be the cathode, and the other electrode will be the anode.

Measuring Standard Reduction Potentials

We can measure the standard reduction potential of a half-cell by constructing a standard cell using a hydrogen reference half-cell and the half-cell whose reduction potential you want to measure. There are two things you need to know: the voltage and the direction of the current. The magnitude of the voltage determines the numerical value of the half-cell potential and the direction of the current determines the sign of the half-cell potential. The cell potential is measured with a voltmeter, which also shows the direction the electrons tend to flow. If $E°_{cell}$ is positive, then the positive terminal on the voltmeter is connected to the cathode and the oxidizing agent at the cathode is stronger than hydrogen ions.

The cell shown in **Figure 7** can be represented as follows:

$$\underset{\text{anode}}{Pt(s) \mid H_2(g) \mid H^+(aq)} \parallel \underset{\text{cathode}}{Cu^{2+}(aq) \mid Cu(s)} \qquad E°_{cell} = +0.34\ V$$

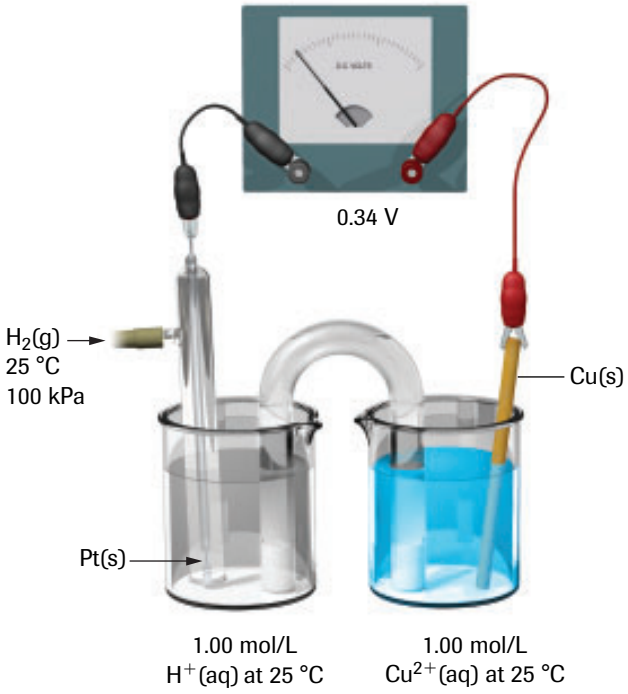


Figure 7
A copper–hydrogen standard cell

cathode: $Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$

anode: $H_2(g) \rightarrow 2\,H^+(aq) + 2\,e^-$

net: $Cu^{2+}(aq) + H_2(g) \rightarrow Cu(s) + 2\,H^+(aq)$

$$E°_{cell} = E°_{r\ cathode} - E°_{r\ anode}$$
$$= 0.34\ V - 0.00\ V$$
$$= +0.34\ V$$

The voltmeter shows that the copper electrode is the cathode and is 0.34 V higher in potential than the platinum anode. If you were to replace the voltmeter with a connecting wire so that the current is allowed to flow, the blue colour of the copper(II) ion disappears and the pH of the hydrogen half-cell decreases as more hydrogen ions are produced and the solution becomes more acidic. Based on this evidence, copper(II) ions are being reduced to copper metal and hydrogen molecules are being oxidized to hydrogen ions. Since this redox reaction is spontaneous, copper(II) ions are stronger oxidizing agents than are hydrogen ions. The standard cell potential, $E°_{cell} = 0.34$ V, is the difference between the reduction potentials of these two half-cells (**Figure 8**).

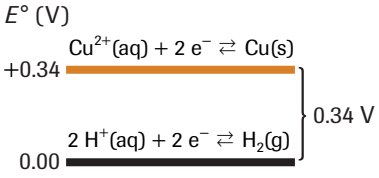


Figure 8
Copper(II) ions are stronger oxidizing agents than hydrogen ions. The cell potential provides a quantitative measurement of how much stronger.

Suppose a standard aluminium half-cell was set up with a standard hydrogen half-cell (**Figure 9**). We can represent the cell as

$$\underset{\text{anode}}{Al(s) \mid Al^{3+}(aq)} \parallel \underset{\text{cathode}}{H^+(aq) \mid H_2(g) \mid Pt(s)} \qquad E°_{cell} = +1.66\ V$$

According to the voltmeter, the platinum electrode is the cathode and the aluminium electrode is the anode. This indicates that hydrogen ions are stronger oxidizing agents than aluminium ions by 1.66 V. Since the reduction potential of hydrogen ions is defined as 0.00 V, the reduction potential of the aluminium ions must be 1.66 V below that of hydrogen, or −1.66 V (**Figure 10**).

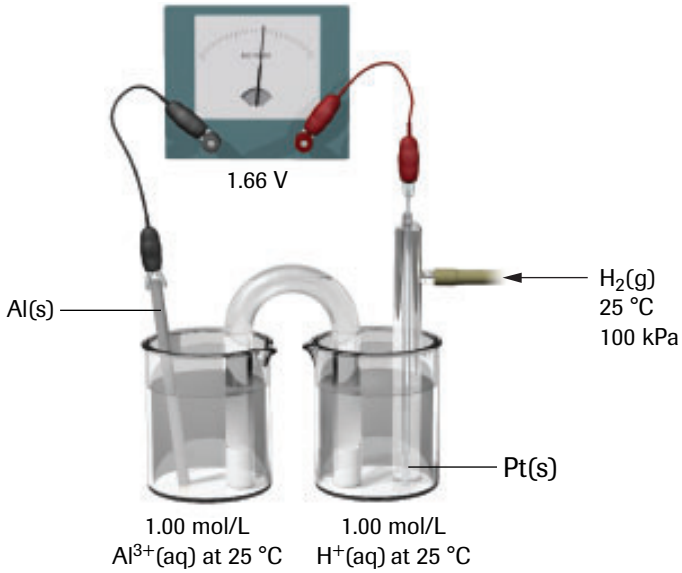


Figure 9
An aluminium–hydrogen standard cell

$$2\,H^+(aq) + 2\,e^- \rightleftarrows H_2(g) \qquad E°_r = 0.00\ V$$
$$Al^{3+}(aq) + 3\,e^- \rightleftarrows Al(s) \qquad E°_r = -1.66\ V$$

The standard cell potential, $E°_{cell} = 1.66$ V, is the difference between the reduction potentials of these two half-cells. To obtain the net or overall cell reaction, add the reduction and oxidation half-reactions, but remember to balance and cancel the electrons.

cathode: $3\,[2\,H^+(aq) + 2\,e^- \rightarrow H_2(g)]$

anode: $2\,[Al(s) \rightarrow Al^{3+}(aq) + 3\,e^-]$

net: $6\,H^+(aq) + 2\,Al(s) \rightarrow 3\,H_2(g) + 2\,Al^{3+}(aq)$

$$E°_{cell} = E°_{r\ \text{cathode}} - E°_{r\ \text{anode}}$$
$$= 0.00\ V - (-1.66\ V)$$
$$= +1.66\ V$$

Notice that the half-reaction equations were multiplied by appropriate factors to balance the electrons, but the *reduction potentials were not altered by the factors used to balance the electrons.* Electric potential represents energy per coulomb of charge (1 V = 1 J/C). Multiplying the aluminium half-reaction by a factor of 2 doubles both the energy and the charge transferred, so that the ratio of energy (J) to charge (C), that is the voltage, is unaffected.

In both of these examples, the strongest oxidizing agent is reduced at the cathode and the strongest reducing agent is oxidized at the anode. The measured cell potential is the difference between the reduction potentials at the cathode and at the anode.

A positive cell potential ($E°_{cell} > 0$) indicates that the net reaction is spontaneous—a requirement for all voltaic cells.

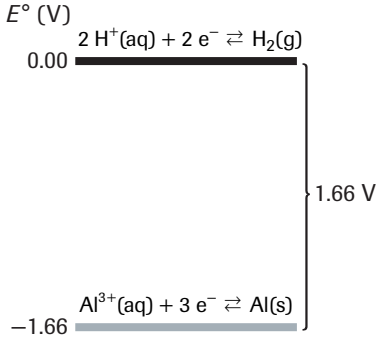


Figure 10
On a redox table, hydrogen ions are stronger oxidizing agents than aluminium ions. The cell potential tells us the hydrogen ions are 1.66 V above aluminium ions.

Learning Tip

Altering the coefficients in a half-reaction equation does **not** affect the reduction potential.

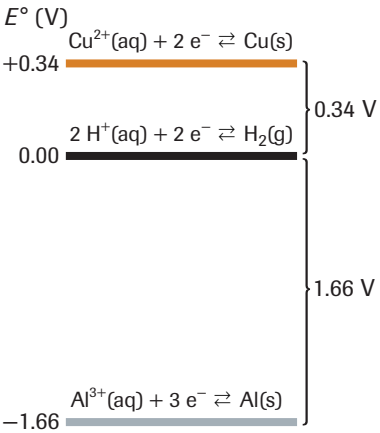


Figure 11
Measurements of standard cell potentials show that the reduction potential of $Cu^{2+}(aq)$ is +0.34 V greater than that of $H^+(aq)$, which is +1.66 V greater than that of $Al^{3+}(aq)$. If you set up a standard cell using copper and aluminium, what would be the cell potential, $E°_{cell}$? (Answer: +2.00 V)

Figure 11 shows the combined results from the copper–hydrogen and aluminium–hydrogen standard cells. This process of measuring standard cell potentials can quickly be extended to more and more oxidizing agents. Notice that this process, although started with the hydrogen reference cell, does not require that it be used for all cell measurements. For example, knowing that the reduction potential of copper(II) ions is 0.34 V, we can now set up many cells that include a standard copper half-cell. A more extensive list of reduction potentials is found in the redox table in Appendix I. You can predict the reaction that occurs spontaneously in any voltaic cell operating under standard conditions using the redox table in Appendix I. The standard cell potential is predicted as follows:

$$E°_{cell} = \underset{\text{cathode}}{E°_r} - \underset{\text{anode}}{E°_r}$$

This order of subtraction is necessary to find out whether the reaction is spontaneous under standard conditions. If $E°_{cell}$ is positive, the reaction is spontaneous.

SUMMARY Rules for Analyzing Standard Cells

You can analyze a standard cell knowing the contents of both half-cells using one or more of the following steps:

- Determine which electrode is the cathode. The cathode is the electrode where the strongest oxidizing agent present in the cell reacts, i.e., the oxidizing agent that is closest to the top on the left side of the redox table. If required, copy the reduction half-reaction for the strongest oxidizing agent and its reduction potential.
- Determine which electrode is the anode. The anode is the electrode where the strongest reducing agent present in the cell reacts, i.e., the reducing agent that is closest to the bottom on the right side of the redox table. If required, copy the oxidation half-reaction (reverse the half-reaction by reading from right to left) for the strongest reducing agent and its reduction potential.
- Determine the overall cell reaction. Balance the electrons for the two half-reaction equations (but do not change the $E°_r$) and add the half-reaction equations.
- Determine the standard cell potential, $E°_{cell}$, using the equation:

$$E°_{cell} = \underset{\text{cathode}}{E°_r} - \underset{\text{anode}}{E°_r}$$

WEB Activity

Simulation—Voltaic Cells Under Standard Conditions

In this computer simulation, you will construct voltaic cells using different half-cells, and then analyze their components and processes. A series of questions will prompt you as you form generalizations about the spontaneity of various reactions.

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COMMUNICATION example 2

A standard lead–dichromate cell is constructed. Write the cell notation, label the electrodes, and calculate the standard cell potential.

Solution

$Pb(s) \mid Pb^{2+}(aq) \,||\, Cr_2O_7^{2-}(aq), H^+(aq), Cr^{3+}(aq) \mid C(s)$

anode — cathode

$E°_{cell} = 1.23\ V - (-0.13\ V)$

$= +1.36\ V$

COMMUNICATION example 3

A standard scandium–copper cell is constructed and the cell potential measured. The voltmeter indicates that copper electrode is positive.

$Sc(s) \mid Sc^{3+}(aq) \,||\, Cu^{2+}(aq) \mid Cu(s) \qquad E°_{cell} = +2.36\ V$

Write and label the half-reaction and net equations, and calculate the standard reduction potential of the scandium ion.

Solution

cathode:	$3\,[Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)]$	$E°_r = +0.34\ V$
anode:	$2\,[Sc(s) \rightarrow Sc^{3+}(aq) + 3\,e^-]$	$E°_r = ?$
net:	$3\,Cu^{2+}(aq) + 2\,Sc(s) \rightarrow 3\,Cu(s) + 2\,Sc^{3+}(aq)$	$E°_{cell} = +2.36\ V$

$2.36\ V = 0.34\ V - E°_r$

$E°_r = -2.02\ V$

Learning Tip

To ensure a correct interpretation, always write the cathode half-reaction of the SOA first. This will help you to remember to subtract the reduction potentials in the correct order.

INVESTIGATION 14.3 *Introduction*

Testing Voltaic Cells

Testing is a procedure that is common to both technology and science. In technology, testing is necessary to determine how a product or process works using criteria such as efficiency, reliability, and cost. In science, testing is a key part in the advancement of knowledge. Scientific concepts are developed and then tested to determine their validity and limitations. New ideas that fail the test then need to be restricted, revised, or replaced.

In your Evaluation, pay particular attention to sources of error or uncertainty, and to limitations of the evidence collected.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the predictions of cell potentials and the identity of the electrodes of various cells.

Problem

In cells constructed from various combinations of copper, aluminium, silver, and zinc half-cells, what are the standard cell potentials, and which is the anode and cathode in each case?

To perform this investigation, turn to page 660.

Practice

10. Standard cells are very important in the scientific study of voltaic cells.
(a) Describe the contents and conditions of a standard cell.
(b) Define the standard cell potential in words and in symbols.

11. For each of the following cells, write the equations for the reactions occurring at the cathode and at the anode, and an equation for the overall or net cell reaction. Calculate the standard cell potential. (Use the redox table in Appendix I.)
(a) $Cr(s) \mid Cr^{2+}(aq) \,||\, Sn^{2+}(aq) \mid Sn(s)$
(b) $Co(s) \mid Co^{2+}(aq) \,||\, SO_4^{2-}(aq), H^+(aq), H_2SO_3(aq) \mid C(s)$
(c) $Pt(s) \mid H_2(g) \mid OH^-(aq) \,||\, OH^-(aq) \mid O_2(g) \mid Pt(s)$

BIOLOGY CONNECTION

Reduction Potentials

Electron transfer in biological systems is governed by the same concepts as electron transfer in simpler chemical systems. In both systems, reduction potentials are used to rank important half-reactions. Cytochromes are proteins that are part of the electron transfer process in cellular respiration; ferredoxins are proteins that are an important part of the electron transfer in photosynthesis. These proteins attach to a metal ion such as Fe^{3+}, greatly influencing its ability to transfer electrons.

(free)
$Fe^{3+}(aq) + e^- \rightleftarrows Fe^{2+}(aq)$
$E°_r = +0.77$ V

(in cytochrome c)
$Fe^{3+}(aq) + e^- \rightleftarrows Fe^{2+}(aq)$
$E°_r = +0.25$ V

(in ferredoxin)
$Fe^{3+}(aq) + e^- \rightleftarrows Fe^{2+}(aq)$
$E°_r = -0.43$ V

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Cell Potentials Under Nonstandard Conditions

The electric potential difference or voltage of a cell decreases slowly as the cell operates. Simultaneously, observable changes such as colour changes and precipitate formation occur. If the cell is left for a very long time, the voltage would eventually become zero and no further changes would be observed in the cell. When people refer to a "dead" cell or battery, this is often what they mean.

The electric potential difference of a cell is a measure of the tendency for electrons to flow. Ideally, during a measurement of the cell potential, a voltmeter should not allow any electrons to flow. If electrons flow, oxidation and reduction reactions occur which, in turn, change the concentrations from the standard 1.0 mol/L value. The value that is measured by a voltmeter represents an electric potential or stored energy, just as the water behind a lock in a canal has gravitational potential energy (**Figure 12(a)**). Connecting the electrodes of a cell in a circuit allows the electrons to flow from the anode to the cathode. This is analogous to opening the outlet and allowing the water to flow from behind the gates to a lower point in front of the gates (**Figure 12(b)**). In both cases, stored potential energy is converted to kinetic energy of electrons or water. If the water available behind a lock is allowed to flow out, then eventually no more water will flow. The level (potential energy) of the water on the two sides of the gate is equalized. An equilibrium is reached with no potential energy difference (**Figure 12(c)**). A similar situation occurs with an operating cell. If electrons are allowed to flow, eventually an equilibrium will be reached when the flow ceases. The rate of the forward reaction, which predominates initially, decreases as the rate of the reverse reaction increases, until the two rates become equal. This is the equilibrium condition and no net flow of electrons will occur. At this time, the electric potential difference as measured by a voltmeter becomes zero.

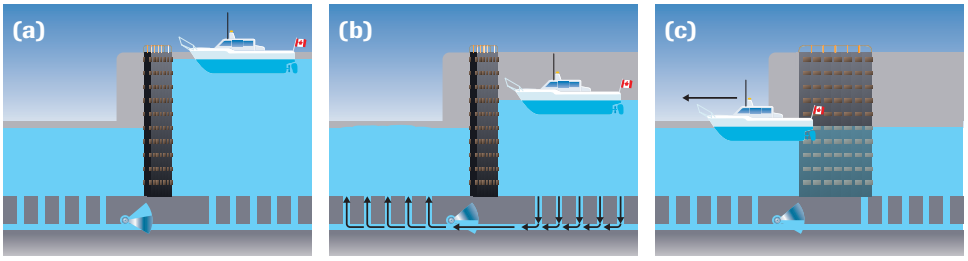

Figure 12

(a) The water behind the gates in a lock has a certain potential energy, ΔE, relative to the bottom of the closed outlet.

(b) When the outlet is opened, water spontaneously flows to the lower level on the other side of the gates. Potential energy, ΔE, is converted to kinetic energy of the flowing water. The water flowing through the outlet is analogous to electron flow.

(c) The flow of water ceases when the levels on both sides of the gates become equal. The gates open, and the ship can then exit to the next lock.

Discrepancies between Measured and Predicted Cell Potentials

Nonstandard conditions for concentrations, temperature, and pressure will cause differences between the cell potentials predicted from a standard redox table and ones measured in a laboratory. However, these differences are generally small if the conditions are relatively close to standard values. Other more important reasons for discrepancies include the purity of the substances, the presence of oxide coatings on metals, and the type and size of the porous boundary.

Practice

12. For each of the following standard cells, refer to the redox table in Appendix I, to represent the cell using the standard cell notation (listing the anode first). Identify the cathode and anode and calculate the standard cell potential without writing half-reaction equations.
(a) copper–lead standard cell
(b) zinc–nickel standard cell
(c) iron(III)–hydrogen standard cell

13. One experimental design for determining the position of a half-cell reaction that is not included in a redox table is shown below. Use the following standard cell, refer to the standard reduction potential of gold in Appendix I, and calculate the reduction potential for the indium(III) ion.

$$\underset{\text{anode}}{In(s) \mid In^{3+}(aq)} \;||\; \underset{\text{cathode}}{Au^{3+}(aq) \mid Au(s)} \qquad E°_{cell} = +1.84\ V$$

14. Any standard half-cell could have been chosen as the reference half-cell, the zero point of the reduction potential scale. What would be the standard reduction potentials for copper and zinc half-cells, assuming that the standard lithium cell were chosen as the reference half-cell with its reduction potential defined as 0.00 V? Justify your answer.

15. List some reasons for differences that might be observed between cell potentials predicted from a table of reduction potentials and cell potentials measured in a laboratory.

16. A zinc–iron cell is constructed and allowed to operate until the measured potential difference becomes zero. What interpretation can be made about the chemical system at this point?

LAB EXERCISE 14.A

Developing a Redox Table

We can use standard cells and their measured cell potentials to develop a redox table.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Purpose

The purpose of this exercise is to use the concepts and rules of standard cells to develop a redox table.

Problem

What is the table of relative strengths of oxidizing and reducing agents based on measured cell potentials?

Design

Several cells are investigated; each cell has at least one half-cell in common with one of the other cells. The cell potentials are measured, and the positive and negative electrodes of each cell are identified.

Evidence

Negative electrode			Positive electrode	
$Pd(s) \mid Pd^{2+}(aq)$	$\|\|$	$Cr_2O_7^{2-}(aq), H^+(aq) \mid C(s)$		$E°_{cell} = +0.28\ V$
$Ti(s) \mid Ti^{2+}(aq)$	$\|\|$	$Tl^+(aq) \mid Tl(s)$		$E°_{cell} = +1.29\ V$
$Tl(s) \mid Tl^+(aq)$	$\|\|$	$Pd^{2+}(aq) \mid Pd(s)$		$E°_{cell} = +1.29\ V$

Figure 13
Large ships have steel hulls. The rusting of steel involves the oxidation of iron in the steel and is a constant headache for shipping companies.

Corrosion

Human history is often divided into different "ages" such as the Copper, Bronze, and Steel ages. These descriptions are based on when these metals were first widely refined and used for tools and weapons. The process of refining a metal is electrochemical in nature and requires energy to recover the pure metal from its naturally occurring compounds (ores). **Corrosion** is also an electrochemical process in which a metal reacts with substances in the environment, returning the metal to an ore-like state. Because we live in an oxidizing (oxygen) environment, oxidation (corrosion) of some metals occur spontaneously. In fact, we need to produce metals such as iron continually to replace the metals lost to corrosion. Preventing corrosion and dealing with the effects of corrosion are major economic and technological problems for our society (**Figure 13**).

As a metal is oxidized, metal atoms lose electrons to form positive ions. A redox table of relative strengths of oxidizing and reducing agents provides the evidence that metals vary greatly in their ability to be oxidized. Some metals, such as gold and silver, are "noble" because they are relatively weak reducing agents. On the other hand, Group 1 and 2 metals are very strong reducing agents and are, therefore, easily oxidized. In general, any metal appearing below the various oxygen half-reactions in a redox table will be oxidized in our environment. Iron (including steel) and aluminium are such metals, and are extensively used as structural materials. Why is the corrosion or rusting of iron such a major problem, but the corrosion of aluminium, which is a much stronger reducing agent, is not? The answer lies primarily in the nature of the oxide that forms on the surface of the metal. A freshly cleaned surface of aluminium rapidly oxidizes in air to form aluminium oxide.

$$4\,Al(s) + 3\,O_2(g) \rightarrow 2\,Al_2O_3(s)$$

The aluminium oxide adheres tightly to the surface of the metal. This prevents further corrosion by effectively sealing any exposed surfaces.

Unfortunately, the iron compounds that form on the surface of exposed iron do not adhere very well. They flake off, exposing new iron to be corroded. In addition, the corrosion of iron is a complex process that is significantly affected by the presence of substances other than oxygen.

Rusting of Iron

Studies of the corrosion of iron have shown that the presence of both oxygen and water is required and the iron is converted into iron hydroxides and oxides. The first step of the mechanism is thought to be the oxidation of iron at a wet exposed surface (**Figure 14**).

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2\,e^-$$

Iron(II) ions diffuse through the water on the iron surface while the electrons easily travel through the iron metal, which is an electrical conductor. The electrons are picked up by oxygen molecules dissolved in water on the surface at a point away from the original oxidation site (**Figure 14**).

$$\tfrac{1}{2}\,O_2(g) + H_2O(l) + 2\,e^- \rightarrow 2\,OH^-(aq)$$

The combination of iron(II) ions and hydroxide ions forms a low-solubility precipitate of iron(II) hydroxide, which is further oxidized by oxygen and water to form iron(III) hydroxide, a yellow-brown solid. The familiar red-brown rust is formed by the dehydration of iron(III) hydroxide to form a mixture of iron(III) hydroxide and hydrated iron(III) oxide. The amount of the hydroxide and the oxide varies, so rust is referred to as a hydrated oxide of indeterminate formula, $Fe_2O_3 \cdot x\,H_2O(s)$.

DID YOU KNOW ?

Rates of Corrosion

A tin can (tin on steel) will corrode completely in about 100 a; an aluminium can in about 400 a; and a glass bottle in about 100 ka. (a, or annum, is the SI unit for year.)

DID YOU KNOW ?

Hydrated Oxide

Iron(III) hydroxide can be converted to iron(III) oxide trihydrate as shown below:

$$2\,Fe(OH)_3(s) \rightarrow Fe_2O_3 \cdot 3\,H_2O(s)$$

It is difficult to determine how much of the iron(III) exists in rust as the hydroxide or hydrated oxide.

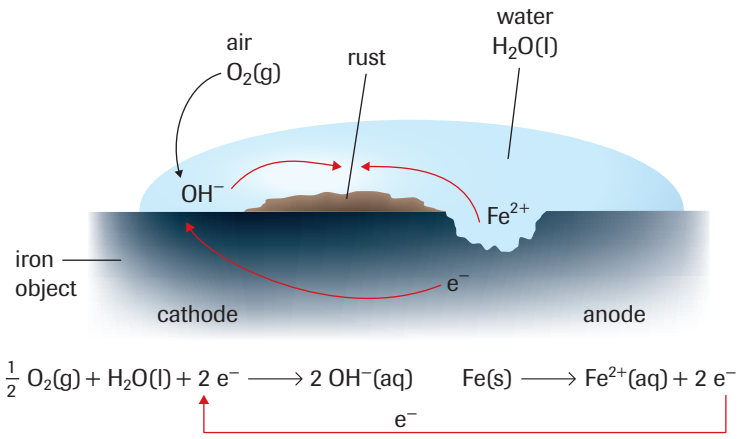

Figure 14
The corrosion of iron is a small electrochemical cell with iron oxidation at one location (the anode) and oxygen reduction at another location (the cathode).

This simplified mechanism for the rusting of iron can be used to explain why certain conditions promote rusting. If the iron is kept in a dry environment (low humidity) or if air has been removed from the water, little or no corrosion occurs (**Figure 15**). Eliminating either water or the oxygen in the water makes the reduction of aqueous oxygen impossible. Iron cannot be oxidized unless a suitable oxidizing agent is present. If oxidizing agents other than oxygen are present, such as certain metal ions, nonmetals, or hydrogen ions, the iron can still be corroded through spontaneous redox reactions. This helps to explain the corrosion of iron in acidic environments; for example, why acid rain corrodes iron faster than natural rain does.

In general, electrolytes accelerate rusting. Ships rust more rapidly in seawater than in fresh water and cars rust more rapidly in places where salt is used on roads. Chloride ions from salt are known to inhibit the adherence of protective oxide coatings on many metals, thus exposing more metal to be corroded. Electrolytes like sodium chloride conduct electricity and improve charge transfer, accelerating the rusting process. Plumbers know that you cannot use steel (iron) straps or nails to hold copper pipes in place because corrosion of the iron will be accelerated. Any moisture that is present sets up an electric cell similar in principle to Volta's original discovery of electricity from dissimilar metals (Section 14.1). As the cell operates, the iron corrodes to form rust. In general, *the rusting of iron requires the presence of oxygen and water and is accelerated by the presence of acidic solutions, electrolytes, mechanical stresses, and contact with less active metals.*

Figure 15
Rusting of exposed iron is almost negligible when the relative humidity is less than 50%. This iron pillar in Delhi, India, has existed for about 1500 years because of the very dry and unpolluted environment.

mini Investigation — Home Corrosion Experiment

Soft drinks are acidic and contain electrolytes. Would different types of soft drink corrode iron at different rates?

Materials: soft drinks (cola, lemon-lime flavoured soda), 2 identical steel nails, 2 plastic glasses

(a) Predict which drink will cause the nails to corrode faster.

- Test your prediction. Place a clean steel nail in a plastic cup filled with the cola (**Figure 16**). Put the other nail in the other cup filled with the other soft drink.
- After at least one day, examine the nails for evidence of any changes.

(b) Record and explain your results.

- Dispose of all materials as directed by your teacher.

Figure 16
What happens to a steel nail in clear soda?

CAREER CONNECTION

Materials Engineering Technologist

How does materials engineering relate to corrosion? Research the entrance requirements, job prospects, and typical salaries for technologists in this field.

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EXTENSION

Corrosion Prevention

You might have seen shiny, reddish copper roofs turn green-grey, after they have been installed for a while. This is a natural patina, that actually helps prevent further corrosion. Scientists are working on developing an artificial patina, that protects the metal much more quickly. Listen to this audio clip to find out more about this process.

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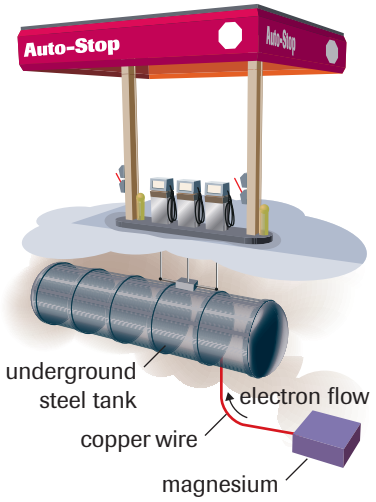

Figure 17
Corrosion of iron involves the oxidation of iron at the anode of a cell. If the iron is attached directly or connected electrically to a metal that is more easily oxidized (a sacrificial anode), then a spontaneous cell develops in which iron is the cathode. The electrolyte of the cell is the moisture in the ground.

Corrosion Prevention

Corrosion is such an important problem in our society that many technologies have been developed, and continue to be developed, to minimize this problem. There are also many career opportunities as engineers or technologists. Methods used for preventing or minimizing the corrosion of iron can be divided into two categories: barrier methods that employ *protective coatings*, and the method of *cathodic protection*. In some critical situations, such as a large fuel tank, both methods may be used.

Protective Coatings

Paint and other similar coatings are a simple method of corrosion prevention. This method works well as long as the surface is completely covered and the coating remains intact. Unfortunately, a scratch or chip in the surface can easily expose a small surface of iron and corrosion begins.

Both tin and zinc are used as metallic coatings. Tin, as in the familiar tin can, adheres well to the iron and provides a strong, shiny coating. The outer surface of the tin coating has a thin, strongly adhering layer of tin oxide that protects the tin as long as the food stored in the can is not too acidic. If a crack or break occurs in the tin layer, moisture can collect in the crack and an electric cell with tin and iron electrodes is established. Since iron is more easily oxidized than tin, iron becomes the anode in this cell. The electrons released by the oxidation of iron flow to the tin and corrosion is accelerated. Evidence of this is the typical iron rust on tin cans that have been crushed and left outside.

A spontaneous electric cell also arises when a zinc coating on an iron object is broken. However, in this case, the zinc is more easily oxidized than the iron. The zinc is preferentially oxidized, preventing corrosion of the iron. Zinc plating (galvanizing) of steel or iron provides double protection: a protective layer and preferential corrosion of the zinc.

Cathodic Protection

According to the redox theory of a cell, oxidation is the loss of electrons and occurs at the anode of a cell. Therefore, an effective method of preventing corrosion of iron is **cathodic protection** in which the iron is forced to become the cathode by supplying the iron with electrons, using either an impressed current or a sacrificial anode.

For a battery or DC generator connected in a circuit, electrons flow out of the negative terminal and into the positive terminal. If the negative terminal is connected to the iron object and the positive terminal is connected to an inert carbon electrode, an electric current is forced to flow to the iron through an electrolyte, such as ground water, from the carbon electrode. The iron is forced to become the cathode and is prevented from corroding. An *impressed current* is an electric current forced to flow toward an iron object by an external potential difference. This method of corrosion prevention requires a constant electric power supply (typically 8 mV) and is used as cathodic protection for pipelines and culverts.

A simpler method of cathodic protection is the use of a sacrificial anode. A *sacrificial anode* is a metal more easily oxidized than iron and connected to the iron object to be protected. The practice of zinc plating (galvanizing) iron objects is a common example of this method. Sacrificial zinc anodes are also connected to the exposed underwater metal surfaces of ships and boats to prevent the corrosion of the iron in the steel. Blocks of magnesium can also be used as sacrificial anodes (**Figure 17**). In all cases, the more active metal (appearing below iron in a half-reaction table) is slowly consumed or sacrificed at the anode, forcing the iron object to be the cathode of the cell.

Practice

17. What are the minimum requirements for the corrosion of iron?

18. List some factors that accelerate or promote the corrosion of iron.

19. Write the balanced net ionic equation for the corrosion of iron to iron(II) ions in the presence of oxygen and water.

20. Although the corrosion of iron is a serious problem, other metals are also corroded in air or other environments. For each of the following situations, use your knowledge of writing and balancing redox equations to write and label the half-reaction and net ionic equations:
(a) Zinc is an active metal that oxidizes when exposed to air and water.
(b) A lead pipe corrodes if it is used to transport acidic solutions that also contain dissolved oxygen.
(c) In dry air, minute quantities of hydrogen sulfide gas can slowly react with silver objects to produce hydrogen gas and silver sulfide, recognized by the dark tarnish on the surface of the silver.

21. You may have noticed that when a car body rusts, the rust appears around the break or chip in the paint, but the damage may extend under the painted surface for some distance.
(a) What is the evidence for damage extending well beyond the break in the paint?
(b) Suggest an explanation why the damage may extend far from the break in the paint.

22. Would a basic solution prevent or slow down the corrosion of iron? Provide your reasoning.

23. Why is a zinc coating on iron better than a tin coating?

24. What are the two methods of cathodic protection and how are they similar?

WEB Activity

Case Study—Galvanizing Steel

A galvanized (zinc-coated) chain-link fence can last a long time without deteriorating (**Figure 18**). In this computer simulation, you will learn about the steps involved in the galvanizing process. After you observe the computer simulations, describe the method and purpose of each of the cleaning and galvanizing steps in the production of galvanized steel. Evaluate the corrosion resistance of a galvanized pipe compared with a painted pipe.

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Figure 18
The metal in this fence is galvanized steel: steel coated with zinc. As the zinc surface layer corrodes, it forms a protective coating that prevents the steel from corroding.

Section 14.2 Questions

1. In the context of a voltaic cell, write a definition of each of the following terms: half-cell, cathode, anode, cation, anion, and inert electrode.

2. State the function of a porous boundary in a voltaic cell and describe two common examples.

3. Distinguish between the external circuit and the internal circuit of a cell.

4. Define a standard cell.

5. (a) For a given cell, how is the cell potential predicted?
(b) What are the restrictions on this prediction?

6. How does the cell potential indicate spontaneity of the reaction?

7. Why are the reactions in voltaic cells always spontaneous? What does this imply about the cell potential?

8. Define the hydrogen reference cell, including contents and conditions. Include the half-cell notation.

9. Why is a reference half-cell necessary?

10. (a) What is the cell potential of a standard cobalt–zinc cell?
(b) What is the theoretical interpretation of this cell potential?
(c) List some factors that may account for the differences you would see between the experimental and predicted values of the cell potential for the cobalt–zinc cell.

11. For each of the following cells
- identify the strongest oxidizing and reducing agents
- write chemical equations to represent the cathode, anode, and overall (net) cell reactions (include the half-cell and cell potentials)
- draw a diagram of each cell, labelling the electrodes, polarity (signs) of electrodes, electrolytes, direction of electron flow, and direction of ion movement

(a) $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

(b) $Sn(s) \mid Sn^{2+}(aq) \parallel Cr_2O_7^{2-}(aq), H^+(aq) \mid C(s)$

12. You can determine a possible identity of an unknown half-cell from the cell potential involving a known half-cell. Use the following evidence and the redox table in Appendix I to determine the reduction potential and possible identity of the unknown $X^{2+}(aq) \mid X(s)$.

$$2\,Ag^+(aq) + X(s) \rightarrow 2\,Ag(s) + X^{2+}(aq) \quad E°_{cell} = +1.08\,V$$

13. An important goal of technology is to provide solutions to practical problems in society. Illustrate this goal by identifying a common problem and the technological solution.

Extension

14. A zinc wire is connected to and buried with a pipeline when it is built (**Figure 19**).

Figure 19
When this pipeline was being constructed, a zinc wire was attached to and buried with the pipe.

(a) Why is this done? Include a brief description of the principles involved.

(b) Is this the only type of corrosion protection used with major pipelines?

(c) Discuss the environmental and safety issues associated with protecting and also not protecting pipelines.

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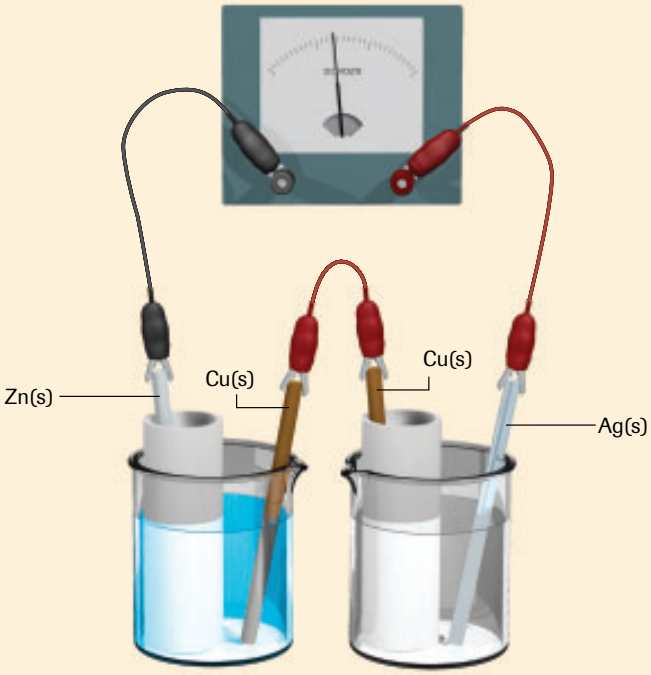
15. Complete the Prediction for the following investigation. Include your reasoning.

Problem
What is the total electric potential difference of two cells connected in series?

Design
Copper–silver and copper–zinc standard cells are connected as shown in **Figure 20**. The total electric potential difference of the two cells is measured with a voltmeter connected to the silver and zinc electrodes.



$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ — $Cu(s) \mid Cu^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

Figure 20
Two standard cells in series

Electrolytic Cells 14.3

Electric cells and batteries used by consumers contain reactants chosen to react spontaneously to convert their chemical energy into electrical energy. These cells or batteries can be used to power a portable music player, start a car, or plate silver metal on jewellery.

A scientific research cell, or voltaic cell, produces electricity spontaneously because each half-cell contains both oxidized and reduced entities. For example,

$$Zn(s) \mid Zn^{2+}(aq) \parallel Pb^{2+}(aq) \mid Pb(s)$$

The cell potential, $E°_{cell}$, is always greater than zero. In a redox table, the strongest oxidizing agent present in the cell will always be above the strongest reducing agent present.

If a cell does not contain all oxidized and reduced species shown in the half-reaction equation, it is possible that the reactants (electrodes and electrolyte) present will not react spontaneously. For example, if lead electrodes are placed in a solution of zinc sulfate and the electrodes are connected with a wire, there is no evidence of any reaction.

$$Pb(s) \mid ZnSO_4(aq) \mid Pb(s)$$

The strongest oxidizing agent present in this cell is $Zn^{2+}(aq)$ and the strongest reducing agent present is Pb(s). A quick check in the redox table shows that the oxidizing agent, $Zn^{2+}(aq)$, is well below the reducing agent, Pb(s), and the $E°_{cell}$ is negative (**Figure 1**).

$Pb^{2+}(aq)$	+	$2\ e^-$	$\rightleftarrows$	Pb(s)
$Ni^{2+}(aq)$	+	$2\ e^-$	$\rightleftarrows$	Ni(s)
$Fe^{2+}(aq)$	+	$2\ e^-$	$\rightleftarrows$	Fe(s)
$Zn^{2+}(aq)$	+	$2\ e^-$	$\rightleftarrows$	Zn(s)

We can calculate the $E°_{cell}$ for the only reaction that could occur.

$$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s) \qquad E°_r = -0.76\ V$$
$$Pb(s) \rightarrow Pb^{2+}(aq) + 2\ e^- \qquad E°_r = -0.13\ V$$
$$Zn^{2+}(aq) + Pb(s) \rightarrow Zn(s) + Pb^{2+}(aq)$$

$$E°_{cell} = \underset{\text{cathode}}{E°_r} - \underset{\text{anode}}{E°_r}$$
$$= -0.76\ V - (-0.13\ V)$$
$$= -0.63\ V$$

Since the $E°_{cell}$ for the reaction is negative, we conclude that the lead will not be oxidized spontaneously in the zinc sulfate solution. Note that the reverse reaction would be spontaneous, but could not occur because neither $Pb^{2+}(aq)$ nor Zn(s) is present initially.

Strictly speaking, the zinc sulfate cell is not a standard cell even if the concentration of the zinc sulfate were 1.0 mol/L because the cell does not contain all entities listed in the half-reaction equations. Therefore, the cell potential that is calculated is not accurate, but will be close enough for our purposes. This cell would not produce electricity because the reaction is nonspontaneous. Why would anyone be interested in a cell like this? Certainly not to use in a battery. However, by supplying electrical energy to a nonspontaneous cell, we can force the reaction to occur. This is especially useful for producing substances, particularly elements. For example, the zinc sulfate cell discussed above is similar to the cell used in the industrial production of zinc metal (**Figure 2**).

Learning Tip

Standard Cells

Recall that a standard cell contains all entities listed in the equation for the half-reaction. In addition, the concentration of aqueous entities is 1.0 mol/L and the conditions are SATP for all substances, including gases at 100 kPa.

Figure 1
According to the redox spontaneity rule, if the strongest oxidizing agent present is below the strongest reducing agent present, no spontaneous reaction will occur ($E°_{cell} < 0$).

Figure 2
Cominco in Trail, British Columbia, operates the world's largest zinc and lead smelter, producing almost 300 kt of zinc annually.

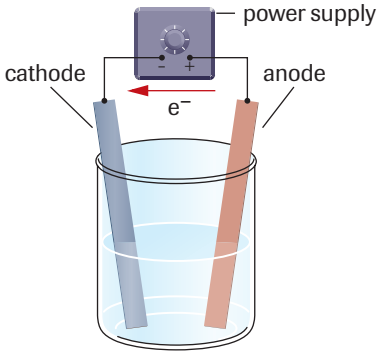


Figure 3
Electrons are pulled from the anode and pushed to the cathode by the battery or power supply.

The term *electrochemical cell* is often used in chemistry to refer to either a cell with a spontaneous reaction, such as electric and voltaic cells, or a cell with a nonspontaneous reaction, which we call an **electrolytic cell** (**Figure 3**). An electrolytic cell consists of a combination of two electrodes, an electrolyte, and an external battery or power source. It uses a process called **electrolysis**, which is the process of supplying electrical energy to force a nonspontaneous redox reaction to occur. The external power supply acts as an "electron pump;" the electric energy is used to do work on the electrons to cause an electron transfer inside the electrolytic cell. In an electrolytic cell, the chemical reaction is the reverse of that of a spontaneous cell. However, most of the scientific principles you have already studied also apply to electrolytic cells (**Table 1**).

$$\text{reactants} \underset{\text{electrolytic cell}}{\overset{\text{electric/voltaic cell}}{\rightleftharpoons}} \text{products} + \text{electrical energy}$$

Table 1 Comparing Electrochemical Cells: Voltaic and Electrolytic

	Voltaic cell	Electrolytic cell
spontaneity	spontaneous reaction	nonspontaneous reaction
standard cell potential, $E°_{cell}$	positive	negative
cathode	• strongest oxidizing agent present undergoes a *reduction* • positive electrode	• strongest oxidizing agent present undergoes a *reduction* • negative electrode
anode	• strongest reducing agent present undergoes an *oxidation* • negative electrode	• strongest reducing agent present undergoes an *oxidation* • positive electrode
direction of electron movement	anode → cathode	anode → cathode
direction of ion movement	anions → anode cations → cathode	anions → anode cations → cathode

Learning Tip

"Positive" and "negative" are labels that are used in many situations, mostly decided by general agreement (convention). For example, "positive" and "negative" can be used to label or describe attitudes, directions, axes on a graph, charges, and electrodes. By convention, a voltaic cell has a cathode labelled positive and an anode labelled negative. In an electrolytic cell, it is customary to reverse these labels. It is best to think of positive and negative for electrodes as labels, not charges.

Secondary Cells: Electric and Electrolytic

A secondary cell is a rechargeable cell such as a nickel–cadmium (Ni–Cd) cell. A secondary cell can be used to illustrate the difference between an electric cell and an electrolytic cell. As the cell discharges, electrical energy is spontaneously produced and the cell functions as an electric cell. When the cell is recharged, the electrical energy forces the products to react and re-form the original reactants. During recharging, the secondary cell is functioning as an electrolytic cell.

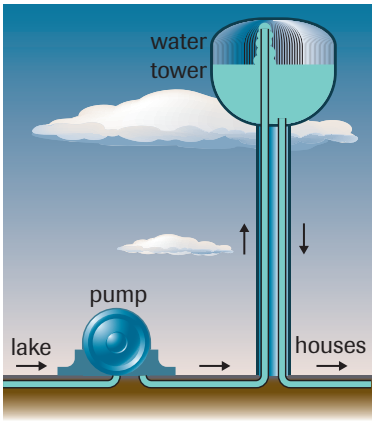


Figure 4
A town's water supply

Practice

1. Describe the type of agent reacting and the process occurring at the cathode and anode of an electrolytic cell.
2. Describe the differences between the cathode and anode of an electrolytic cell and a voltaic cell.
3. Describe the direction of movement of electrons and ions within an electrolytic cell.
4. In a town's water supply (**Figure 4**), water is pumped from a lake into a water tower, which supplies water to the town's houses. Describe how this is an analogy for a secondary cell.

INVESTIGATION 14.4 *Introduction*

A Potassium Iodide Electrolytic Cell

Electrolytic cells were discovered before the science was understood. However, as with all successful technological inventions, the important criteria was that it worked, not why it worked. Eventually, chemists understood the science and were able to explain why electrolytic cells work.

In the Evaluation, suggest changes to the Design, Materials, and Procedure that would improve the Evidence.

Purpose

The purpose of this investigation is to use diagnostic tests to determine the reaction products of an electrolytic cell.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 3)
○ Hypothesis	○ Procedure	
○ Prediction	● Evidence	

Problem

What are the products of the reaction during the operation of an aqueous potassium iodide electrolytic cell?

Design

Inert electrodes are placed in a 0.50 mol/L solution of potassium iodide, and a battery or power supply provides a direct current of electricity to the cell. The litmus and halogen diagnostic tests are conducted to test the solution near each electrode before and after the reaction.

To perform this investigation, turn to page 661.

The Potassium Iodide Electrolytic Cell: A Synthesis

In the potassium iodide electrolytic cell (**Figure 5**), litmus paper does not change colour in the initial solution and turns blue only near the electrode from which gas bubbles. At the other electrode, a yellow-brown colour and a dark precipitate forms. The yellow-brown substance produces a purplish-red colour in a halogen test. This chemical evidence agrees with the interpretation supplied by the following half-reaction equations. According to the redox table of relative strengths of oxidizing and reducing agents, water is the stronger oxidizing agent present and iodide ions are the stronger reducing agents present in a potassium iodide solution.

	OA		SOA
	$K^+(aq)$	$I^-(aq)$	$H_2O(l)$
		SRA	RA

cathode: $2\,H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2\,OH^-(aq)$
gas bubbles; blue litmus

anode: $2\,I^-(aq) \rightarrow I_2(s) + 2\,e^-$
purplish-red in hexane

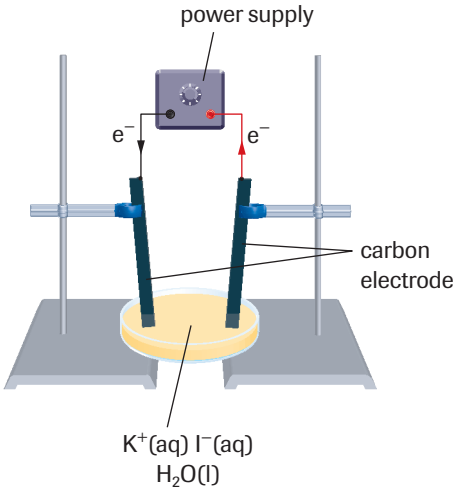


Figure 5
A power supply provides the energy for the chemical reactions at the two electrodes.

Evidence from the study of this and many other aqueous electrolytic cells suggests that the generalizations for voltaic cells also apply to electrolytic cells. From a theoretical perspective, the strongest oxidizing agent present in a particular mixture has the greatest attraction for electrons and gains electrons at the cathode. Notice that it does not matter where the electrons originate, from a power supply or directly from another electrode. The strongest reducing agent present in the mixture has the least attraction for electrons and loses electrons at the anode. In other words, the theoretical definitions of cathode and anode are the same for both voltaic and electrolytic cells (**Table 1**).

Observation of a potassium iodide cell indicates that the transfer of electrons is not spontaneous. When a voltage is supplied to the cell, electrons flow from the negative terminal of the battery toward the cathode of the electrolytic cell and are consumed (gained) by water molecules, which have the more positive reduction potential. Simultaneously, electrons flow from iodide ions on the surface of the anode to the positive terminal of the battery. This explanation agrees with previous redox concepts and with the observations, so we can accept the explanation. Predictions of cathode, anode, and overall cell reactions for electrolytic cells follow the same steps outlined for voltaic cells in Section 14.2.

Learning Tip

At first glance, the cell notation, C(s) | KI(aq) | C(s), does not show that water is involved in the reaction. Of course, the subscript (aq) means "dissolved in water." Don't forget to consider the presence of water because it is often a reactant in aqueous electrolytic cells.

SAMPLE problem 14.1

What are the cell reactions and the cell potential of the aqueous potassium iodide electrolytic cell?

First, we identify the major entities in the solution and use the redox table in Appendix I to identify the strongest oxidizing and reducing agents, as we did in Chapter 13.

OA SOA
$K^+(aq)$, $I^-(aq)$, $H_2O(l)$
SRA RA

Now we can write the half-reaction equations and calculate the cell potential. The potassium iodide cell is not a standard cell because the products of the reactions are not present initially. Therefore, the reduction potentials given in the table of half-reactions are not strictly applicable, but we will use them to approximate the cell potential.

cathode: $2\,H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2\,OH^-(aq)$ $\quad E°_r = -0.83\ V$

anode: $2\,I^-(aq) \rightarrow I_2(s) + 2\,e^-$ $\quad E°_r = +0.54\ V$

net: $2\,H_2O(l) + 2\,I^-(aq) \rightarrow H_2(g) + 2\,OH^-(aq) + I_2(s)$

$$E°_{cell} = \underset{\text{cathode}}{E°_r} - \underset{\text{anode}}{E°_r}$$
$$= -0.83\ V - (+0.54\ V)$$
$$= -1.37\ V$$

A negative sign for a cell potential indicates that the chemical process is nonspontaneous. The more negative the cell potential, the more energy is required. In Sample Problem 14.1, electrons must be supplied with a minimum of +1.37 V from an external battery or other power supply to force the cell reactions. In practice, however, a greater voltage is required, for example, to make the reaction occur at a reasonable rate.

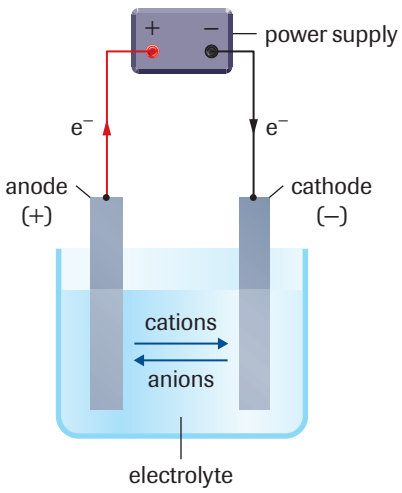


Figure 6
A generic electrolytic cell

SUMMARY Procedure for Analyzing Electrolytic Cells

The procedure for analyzing electrolytic cells is essentially the same as for voltaic cells.

- Use the redox table (Appendix I) to identify the strongest oxidizing and reducing agents present. (Do not forget to consider water for aqueous electrolytes.)
- Write equations for the reduction (cathode) and oxidation (anode) half-reactions. Include the reduction potentials if required.
- Balance the electrons and write the net cell reaction including the cell potential.

$$E°_{cell} = \underset{\text{cathode}}{E°_r} - \underset{\text{anode}}{E°_r}$$

- If required, state the minimum electric potential (voltage) to force the reaction to occur. (The minimum voltage is the absolute value of $E°_{cell}$.)
- If a diagram is requested, use the general outline shown in **Figure 6** and add specific labels for chemical entities.

COMMUNICATION example 1

An electrolytic cell containing cobalt(II) chloride solution and lead electrodes is assembled. The notation for the cell is as follows:

$$Pb(s) \mid Co^{2+}(aq), Cl^{-}(aq) \mid Pb(s)$$

(a) Predict the reactions at the cathode and anode, and in the overall cell.
(b) Draw and label a cell diagram for this electrolytic cell, including the power supply.
(c) What minimum voltage must be applied to make this cell work?

Solution

(a) SRA SOA

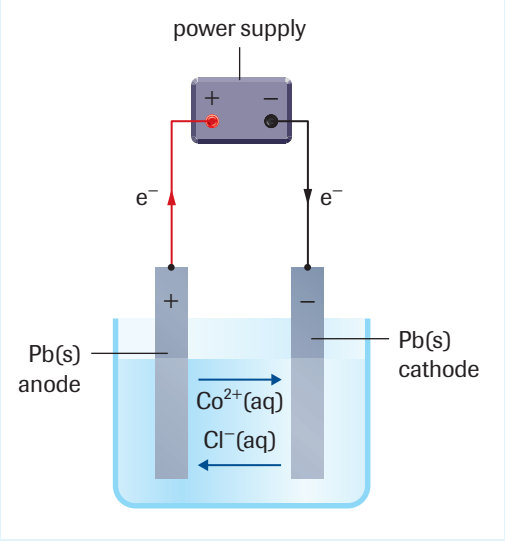
$$Pb(s) \mid Co^{2+}(aq), Cl^{-}(aq) \mid Pb(s)$$

cathode: $Co^{2+}(aq) + 2\,e^{-} \rightarrow Co(s)$

anode: $Pb(s) \rightarrow Pb^{2+}(aq) + 2\,e^{-}$

net: $Co^{2+}(aq) + Pb(s) \rightarrow Co(s) + Pb^{2+}(aq)$

(b)



(c) $E^{\circ}_{cell} = E^{\circ}_{r\ cathode} - E^{\circ}_{r\ anode}$

$= -0.28\ V - (-0.13\ V)$

$= -0.15\ V$

According to the redox table, a minimum voltage of 0.15 V is required.

DID YOU KNOW ?

Hydrometallurgy

Hydrometallurgy is the process of extracting metals or their compounds from ores using aqueous solutions. Sherritt Metals of Fort Saskatchewan, Alberta is a world leader in high pressure hydrometallurgy of nickel and cobalt ores. Sherritt uses an ammonia leaching of nickel and cobalt sulfides followed by a chemical reduction using hydrogen as a reducing agent. Other hydrometallurgical operations use electrolysis instead of a chemical reduction to obtain or purify metals.

COMMUNICATION example 2

An electrolytic cell is set up with a power supply connected to two nickel electrodes immersed in an aqueous solution containing cadmium nitrate and zinc nitrate. Predict the equations for the initial reaction at each electrode and the net cell reaction. Calculate the minimum voltage that must be applied to make the reaction occur.

Solution

SRA SOA

$Ni(s), H_2O(l), Cd^{2+}(aq), NO_3^-(aq), Zn^{2+}(aq)$

cathode:	$Cd^{2+}(aq) + 2\,e^- \rightarrow Cd(s)$	$E^\circ_r = -0.40$ V
anode:	$Ni(s) \rightarrow Ni^{2+}(aq) + 2\,e^-$	$E^\circ_r = -0.26$ V
net:	$Cd^{2+}(aq) + Ni(s) \rightarrow Cd(s) + Ni^{2+}(aq)$	

$$E^\circ_{cell} = \underset{\text{cathode}}{E^\circ_r} - \underset{\text{anode}}{E^\circ_r}$$
$$= -0.40\text{ V} - (-0.26\text{ V})$$
$$= -0.14\text{ V}$$

According to the redox table, a minimum voltage of 0.14 V is required.

Practice

5. Predict the cathode, anode, and net cell reactions for each of the following electrolytic cells. Calculate the minimum potential difference that must be applied to force the cell reaction to occur.
(a) $C(s) \mid Ni^{2+}(aq), I^-(aq) \mid C(s)$ (b) $Pt(s) \mid Na^+(aq), OH^-(aq) \mid Pt(s)$

6. What is the minimum electric potential difference of an external power supply that produces chemical changes in the following electrolytic cells?
(a) $C(s) \mid Cr^{3+}(aq), Br^-(aq) \mid C(s)$ (b) $Cu(s) \mid Cu^{2+}(aq), SO_4^{2-}(aq) \mid Cu(s)$

INVESTIGATION 14.5 *Introduction*

Electrolysis (Demonstration)

Scientific knowledge progresses by the experimental testing of ideas. The more rigorous the test, the more certain the knowledge or the better the chance of making new discoveries.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	○ Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this demonstration is to test the method of predicting the products of electrolytic cells.

Problem

What are the products of electrolytic cells containing one of the following aqueous solutions: copper(II) sulfate, sodium sulfate, and sodium chloride?

Design

The electrolysis of the aqueous copper(II) sulfate is carried out in a U-tube, and the electrolysis of aqueous sodium sulfate and sodium chloride is carried out in a Hoffman apparatus so that any gases produced can be collected. Diagnostic tests with necessary control tests (before electrolysis) are conducted to determine the presence of the predicted products.

To perform this investigation, turn to page 662.

Evaluation of Predicted Reactions—The Chloride Anomaly

As you know from Investigation 14.5, some redox reactions predicted using the strongest oxidizing and reducing agents from a redox table do not always occur in an electrolytic cell. Like other chemical processes, a half-cell reaction at an electrode has an activation energy (Unit 6) that varies for different half-reactions and conditions. Therefore, the *actual* reduction potential required for a particular half-reaction and the *reported* half-reaction reduction potential may be quite different. This difference is known as the half-cell overvoltage. It is generally much greater for the production of oxygen than for the production of chlorine, for example, in the electrolysis of aqueous chloride compounds. The explanation for overvoltage is well beyond the level of this book. *As an empirical rule, you should recognize that chlorine gas is produced instead of oxygen gas in situations where chloride and water are the only reducing agents present.*

Learning Tip

There are exceptions to all rules and generalizations. You only need to remember the chloride anomaly. This occurs during the electrolysis of solutions containing the chloride ion. Since water is the strongest reducing agent present, water should react at the anode. However, the chloride ions react preferentially to water molecules.

SUMMARY Electrolytic Cells

- An electrolytic cell is based upon a reaction that is nonspontaneous; the $E°_{cell}$ for the reaction is negative. An applied voltage of at least the absolute value of $E°_{cell}$ is required to force the reactions to occur.
- The strongest oxidizing agent undergoes reduction at the cathode (negative electrode).
- The strongest reducing agent undergoes oxidation at the anode (positive electrode).
- Electrons are forced by a power supply to travel from the anode to the cathode through the external circuit.
- Internally, anions move toward the anode and cations move toward the cathode.

Practice

7. List the main similarities between a voltaic cell and an electrolytic cell.

8. What is the key difference between voltaic and electrolytic cells?

9. Why is the procedure for analyzing voltaic and electrolytic cells so similar?

10. Explain why a power supply is necessary for an electrolytic cell.

11. Which of the following cells would produce a spontaneous reaction? Justify each answer, using the cell potential.
(a) $C(s) \mid Cr(NO_3)_2(aq) \mid C(s)$ (b) $Cu(s) \mid FeCl_3(aq) \mid Cu(s)$

12. For each of the following electrolytic cells, write equations for the cathode and anode half-reactions and the net reaction. Determine the minimum potential difference that must be applied to make the cell operate.
(a) $C(s) \mid K_2SO_4(aq) \mid Cd(s)$ (b) $Pt(s) \mid SnBr_2(aq) \mid Pt(s)$

13. Draw a diagram of an electrolytic cell containing a zinc iodide solution and inert carbon electrodes.
- Label the power supply and electrodes, including signs, the electrolyte, and the directions of electron and ion movements.
- Write half-reaction and net equations.
- Calculate the cell potential, using standard values.

14. State the chloride anomaly, and include how it is recognized.

15. Describe a specific consumer product that you use sometimes as an electric cell and sometimes as an electrolytic cell. What practical problem does this technology solve?

Science and Technology of Electrolysis

Volta's invention of the electric cell in 1800 immediately resulted in many discoveries in chemistry. One discovery was that electric cells could be used as an electric power source for electrolytic cells. Many natural substances, such as soda (sodium carbonate) and potash (potassium carbonate) that were thought to be elements, were shown to be composed of the previously unknown elements sodium and potassium. Industrial applications of electrolytic cells include the production of elements, the refining of metals, and the plating of metals onto the surface of an object. The study of electrolysis in industry reveals the strong relationship between science and technology.

Production of Elements

Most elements occur naturally combined with other elements in compounds. For example, ionic compounds of sodium, potassium, lithium, magnesium, calcium, and aluminium are abundant, but these reactive metals are not found uncombined in nature. The explanation is that the reduction potentials for these metals are very negative. Consequently, the metals are easily oxidized by practically all other substances. Even water has a more positive reduction potential than any of these metal ions. If the metals did exist naturally, a spontaneous reaction would convert them into their ions.

Figure 7
In his youth, Sir Humphry Davy (1778–1829) worked as an assistant to a physician who was interested in the therapeutic properties of gases. Davy studied nitrous oxide (laughing gas) by conducting experiments on himself. He was eventually fired from his job, supposedly because of his liking for explosive chemical reactions. Davy's main fame came from his experiments with electricity. He constructed a voltaic pile with over 250 metal plates. He used this powerful cell to decompose stable compounds and discovered the elements sodium, potassium, barium, strontium, calcium, and magnesium. Given his habit of tasting, inhaling, and exploding new chemicals, it is not surprising that he was an invalid in his early thirties and died in middle age, probably of chemical poisoning.

SOA

$$Zn^{2+}(aq) + 2\,e^- \rightleftarrows Zn(s)$$
$$2\,H_2O(l) + 2\,e^- \rightleftarrows H_2(g) + 2\,OH^-(aq)$$
$$Mg^{2+}(aq) + 2\,e^- \rightleftarrows Mg(s)$$
$$Na^+(aq) + e^- \rightleftarrows Na(s)$$

Many metals can be produced by electrolysis of solutions of their ionic compounds, but two difficulties arise. First, many naturally occurring ionic compounds have a low solubility in water and second, water is a stronger oxidizing agent than many active metal cations. To overcome these difficulties, we can use a technological design in which water is not present. Fortunately, ionic compounds can be melted. These molten ionic compounds are good electrical conductors and can function as the electrolyte in a cell.

The production of active metals (strong reducing agents) from their minerals typically involves the electrolysis of molten compounds of the metal, a technology first used in the scientific work of Humphry Davy (**Figure 7**). Strontium metal was one of many active metals discovered by Davy using the electrolysis of molten salts. Strontium chloride was first melted in an electrolytic cell with inert electrodes. In this cell, there are only two kinds of ions present, $Sr^{2+}(l)$ and $Cl^-(l)$. You may recall from the previous chapter that metal cations generally tend to undergo a reduction and nonmetal anions tend to undergo an oxidation. In this cell, there are no other competing substances. Therefore, the strontium ions will consume (gain) electrons at the cathode to form strontium metal:

$$Sr^{2+}(l) + 2\,e^- \rightarrow Sr(s) \qquad \text{(reduction at the cathode)}$$

At the anode, chloride ions will give up (lose) electrons to form chlorine gas:

$$2\,Cl^-(l) \rightarrow Cl_2(g) + 2\,e^- \qquad \text{(oxidation at the anode)}$$

The electrons are balanced. Adding the two equations gives the overall reaction in the cell.

$$Sr^{2+}(l) + 2\,Cl^-(l) \rightarrow Sr(s) + Cl_2(g)$$

This reaction would not be possible in an aqueous solution because water is a stronger oxidizing agent (has a more positive reduction potential) than aqueous strontium ions.

In molten-salt electrolysis, metal cations move to the cathode and are reduced to metals, and nonmetal anions move to the anode and are oxidized to nonmetals.

COMMUNICATION example 3

Lithium is the least dense of all metals and is a very strong reducing agent; both qualities make it an excellent anode for batteries. Lithium can be produced by electrolysis of molten lithium chloride at a temperature greater than 605 °C, the melting point of lithium chloride. Write the equations for the cathode and anode half-reactions, and the net cell reaction.

Solution

cathode: $2\,[Li^+(l) + e^- \rightarrow Li(s)]$

anode: $2\,Cl^-(l) \rightarrow Cl_2(g) + 2\,e^-$

net: $2\,Li^+(l) + 2\,Cl^-(l) \rightarrow 2\,Li(s) + Cl_2(g)$

Learning Tip

No reduction potentials can be listed for the electrolysis of a molten salt. The redox table in Appendix I lists only electric potentials for half-reactions in 1.0 mol/L aqueous solutions at SATP.

Production of Aluminium

Aluminium is the third most abundant element on Earth. It was discovered in France in the early 1800s. At the time, aluminium was more expensive than gold. The wonderful properties of aluminium—shiny, light, strong, and corrosion resistant—made it ideal for jewellery and cutlery, so there was a high demand for the metal, especially among the aristocracy. However, the supply of aluminium was limited because the technology for producing aluminium was not yet practical or economically viable for mass production.

Initial efforts to produce aluminium by electrolysis were unproductive because its common ore, $Al_2O_3(s)$, has a high melting point, 2072 °C. No material could be found to hold the molten compound. In 1886, two scientists working independently and knowing nothing of each other's work made the same discovery. Charles Hall in the United States and Paul Héroult in France discovered that $Al_2O_3(s)$ dissolves in a molten mineral called cryolite, Na_3AlF_6. In this design, the cryolite acts as an inert solvent for the electrolysis of aluminium oxide and forms a molten conducting mixture with a melting point around 1000 °C. Aluminium (m.p. 660 °C) can be produced electrolytically from this molten mixture (**Figure 8**). This discovery had an immediate effect on the supply and cost of aluminium. Around 1855, aluminium was sold for $45 000 per kilogram; a few years after the Hall–Héroult invention, the cost was about 90 cents per kilogram.

DID YOU KNOW?

Aluminium Production in Canada

The production of aluminium is important to Canada's economy, although Canada does not have large deposits of aluminium ore. Hydroelectric power is used to produce aluminium metal from concentrated, imported bauxite in an electrolytic cell. Recycling aluminium from soft drink and beer cans requires only 5% of the energy required to produce aluminium by electrolysis.

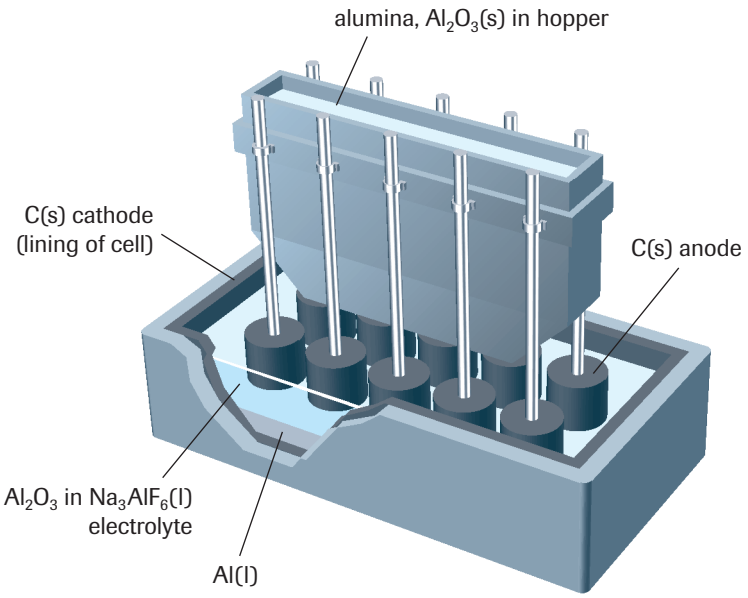


Figure 8
The Hall–Héroult cell for the production of aluminium. The cathode is the carbon lining of the steel cell. At the cathode, aluminium ions are reduced to produce liquid aluminium, which collects at the bottom of the cell and is periodically drained away. At the carbon anode, oxide ions are oxidized to produce oxygen gas. The oxygen produced at the anode reacts with the carbon electrodes, producing carbon dioxide, so these electrodes must be replaced frequently.

EXTENSION

CBC radio ONE

QUIRKS & QUARKS

Aluminium Energy Source

Laptops and even some cars of the future may use aluminium as an energy source. It takes a great deal of electrical energy to produce aluminium and some of this energy can be recovered. Two University of British Columbia researchers have devised a way to produce hydrogen from aluminium using only salt and water. The hydrogen that is produced can be used to power a fuel cell for portable electronics or even a car.

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Figure 9
Large salt deposits sit hundreds of metres below Fort Saskatchewan, Alberta. This means that Fort Sakatchewan is a good location for Dow's chlor–alkali plant because the main raw material is available on site.

Aluminium oxide is obtained from bauxite, an aluminium ore. Once the ore is purified, the aluminium oxide is dissolved in molten cryolite and it dissociates into individual ions. The reactions occurring at the electrodes in a Hall–Héroult cell are summarized below:

cathode: $4[Al^{3+}(\text{cryolite}) + 3\,e^- \rightarrow Al(l)]$

anode: $3[2\,O^{2-}(\text{cryolite}) \rightarrow O_2(g) + 4\,e^-]$

$$4\,Al^{3+}(\text{cryolite}) + 6\,O^{2-}(\text{cryolite}) \rightarrow 4\,Al(l) + 3\,O_2(g)$$

The overall cell reaction is a decomposition of aluminium oxide.

$$2\,Al_2O_3(s) \rightarrow 4\,Al(s) + 3\,O_2(g)$$

The Chlor–Alkali Process

The most important nonmetal produced by electrolysis is chlorine. More than 95% of the world production of chlorine and almost 100% of the world production of sodium hydroxide is done using the chlor–alkali process, which is a reaction that you studied in Investigation 14.5. The chlor–alkali process is the electrolysis of aqueous sodium chloride (brine) to produce chlorine, hydrogen, and sodium hydroxide:

$$2\,NaCl(aq) + 2\,H_2O(l) \rightarrow \underset{\text{(anode product)}}{Cl_2(g)} + \underset{\text{(cathode products)}}{H_2(g)} + 2\,NaOH(aq)$$

You may recall that the anode reaction in this electrolysis was unexpected in Investigation 14.5 because water is a slightly stronger reducing agent than chloride ions. This is the chloride anomaly and is an important exception to the rules for predicting half-reactions.

Dow Chemical in Fort Saskatchewan, Alberta, has an ideal location for its chlor–alkali plant (**Figure 9**). The sodium chloride is extracted from large underground deposits using hot water and pumped to electrolytic cells at the surface. Dow uses diaphragm electrolytic cells, but newer membrane electrolytic cells are becoming more common in the chlor–alkali industry. These two technologies are similar, but the membrane cell has the advantage of producing sodium hydroxide with very little sodium chloride contamination. This is accomplished by using an ion-exchange membrane that allows sodium ions, but not chloride ions, to move from the anode to the cathode compartments (**Figure 10**).

Hydrogen gas is used to make ammonia, hydrogen peroxide, and margarine, and to crack petroleum. It may also be used on site as a fuel to produce electricity. Chlorine is used as a disinfectant for drinking water and to manufacture bleach (sodium hypochlorite), plastics, pesticides, and solvents. Sodium hydroxide is used on a large scale in industry to make cellophane, pulp and paper, aluminium, and detergents.

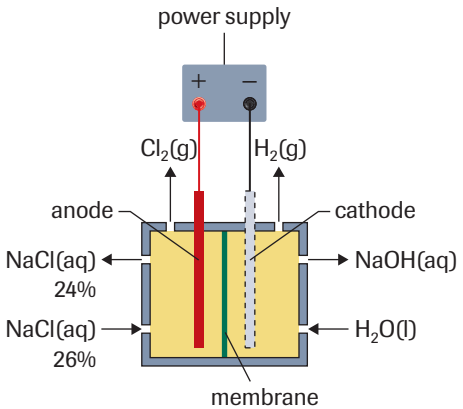

Figure 10
Membrane cells are a relatively recent technology. Most companies throughout the world are switching to these cells from the older diaphragm and mercury cells.

Practice

16. Describe two difficulties associated with the electrolysis of aqueous ionic compounds in the production of active metals. What is a technological solution that overcomes both difficulties?

17. Scandium is a metal with a low density and a melting point that is higher than that of aluminium. These properties are of interest to engineers who design space vehicles. Scandium metal is produced by electrolysis of molten scandium chloride. List all ions present in the electrolysis cell, and write the equations for the reactions that occur at the cathode and anode, and the net cell reaction.

18. Write the cathode, anode, and cell reaction equations for the chlor–alkali process.

19. The following statements summarize the steps in the chemical technology of obtaining magnesium from seawater. Write a balanced equation to represent each reaction.

(a) Slaked lime (solid calcium hydroxide) is added to seawater (ignore all solutes except $MgCl_2(aq)$) in a double displacement reaction to precipitate magnesium hydroxide.

(b) Hydrochloric acid is added to the magnesium hydroxide precipitate.

(c) After the magnesium chloride product is separated and dried, it is melted in preparation for electrolysis. List all ions present in the electrolysis, and write the equations for the reactions that occur at the cathode and anode, and the net cell reaction.

(d) An alternative process produces magnesium from dolomite, a mineral containing $CaCO_3$ and $MgCO_3$. Suggest some technological advantages and disadvantages of the dolomite process compared with the seawater process.

20. What products in your home may have originated from substances produced in the chlor–alkali process?

21. Why should we recycle metals such as aluminium? State several arguments that you might use in a debate.

Refining of Metals

In the production of metals, the initial product is usually an impure metal. Impurities are often other metals that come from various compounds in the original ore. To purify or refine a metal, a variety of methods are used. However, a common method, known as *electrorefining*, uses an electrolytic cell to obtain high-grade metals at the cathode from an impure metal at the anode.

A good example is the electrorefining of copper. The presence of impurities in copper lowers its electrical conductivity, not a desirable property considering that one of the most common uses of copper is in electrical wiring. The initial smelting process produces copper that is about 99% pure, containing some silver, gold, platinum, iron, and zinc. These valuable impurities can be recovered and sold to help pay for the process. As shown in **Figure 11**, a slab of impure copper is the anode of an electrolytic cell that contains copper(II) sulfate dissolved in sulfuric acid. The cathode is a thin sheet of very pure copper. As the cell operates, copper and some of the other metals in the anode are oxidized, but only copper is reduced at the cathode. An understanding of oxidation, reduction, and reduction potentials allows precise control over what is oxidized and what is reduced; after electrorefining, the copper is about 99.98% pure. The half-reactions are:

cathode:	reduction of copper	$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$
anode:	oxidation of copper	$Cu(s) \rightarrow Cu^{2+}(aq) + 2\,e^-$
	oxidation of zinc	$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$
	oxidation of iron	$Fe(s) \rightarrow Fe^{2+}(aq) + 2\,e^-$

DID YOU KNOW?

The Pidgeon Process

Lloyd Pidgeon (1903–1999) was born in Markham, Ontario, studied undergraduate chemistry at the University of Manitoba, and obtained his Ph.D. from McGill University in 1929. Later, while working at the National Research Council in Ottawa, Pidgeon developed the first process for producing high-quality magnesium metal from dolomite (calcium magnesium carbonate). This led to the formation of Dominion Magnesium Ltd. using dolomite mined in the Ottawa Valley. The Pidgeon Process is still used in many countries to produce magnesium.

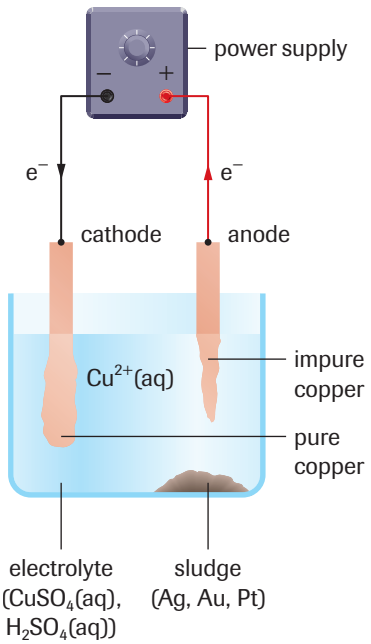

Figure 11
Only copper and metals more easily oxidized than copper, such as iron and zinc, are oxidized to ions and dissolve at the anode. Only copper is reduced at the cathode. Other impurities in the anode, such as silver, gold, and platinum, do not react; these fall to the bottom of the cell as a sludge called anode mud.

Another method of purifying metals is to reduce metal cations from a molten or aqueous electrolyte at the cathode of an electrolytic cell, much like the production of elements discussed previously. This method is known as *electrowinning*. Using a molten salt is the only way to obtain some active metals, such as those in Group 1. Many other metals, such as zinc, can be produced by electrowinning an aqueous solution. For example, Cominco's operation at Trail, British Columbia, uses the electrolysis of an acidic zinc sulfate solution with a specially treated lead anode to deposit very pure zinc metal at the cathode.

cathode: $2[Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)]$

anode: $2\,H_2O(l) \rightarrow O_2(g) + 4\,H^+(aq) + 4\,e^-$

net: $2\,Zn^{2+}(aq) + 2\,H_2O(l) \rightarrow 2\,Zn(s) + O_2(g) + 4\,H^+(aq)$

Electroplating

Several metals, such as silver, gold, zinc, and chromium, are valuable because of their resistance to corrosion. However, products made from these metals in their pure form are either too expensive or they lack suitable mechanical properties, such as strength and hardness. To achieve the best compromise among price, mechanical properties, appearance, and corrosion resistance, utensils and jewellery may be made of a relatively inexpensive, yet strong alloy such as steel, and then coated (plated) with another metal or alloy to improve appearance or corrosion resistance. Plating of a metal at the cathode of an electrolytic cell is called *electroplating* and is a common technology that is used to cover the surface of an object with a thin layer of metal.

The process for plating metals is obtained by systematic trial and error, involving the careful manipulation of one possible variable at a time. In this situation, a scientific perspective helps identify variables, but cannot usually provide successful predictions.

The development and use of electric cells preceded scientific understanding of the processes involved. Today, we still have examples of technological processes that are not fully understood, such as chromium plating (**Figure 12**) and silver plating. For example, there is no satisfactory explanation for why silver deposited during electrolysis of a silver nitrate solution does not adhere well to any surface, whereas silver plated from silver cyanide solution does.

Electroplating is only one method used to cover the surface of an object with a metal. One other method is vapour deposition, in which metal vapour is condensed on the surface. Another method is dipping, in which an object is dipped into a molten metal that solidifies on the surface. Zinc-plated nails for exterior use are made by dipping.

CAREER CONNECTION

Chemical Technologist

Chemical technologists may work closely with scientists and engineers studying electroplating processes. What education and training is required for this job? Outline some current job opportunities in this area, including typical salaries.

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Figure 12
Chromium is best plated from a solution of chromic acid. A thin layer of chromium metal is very shiny and, like aluminium, protects itself from corrosion by forming a tough oxide layer.

SUMMARY Applications of Electrolytic Cells

- In molten-salt electrolysis, metal cations are reduced to metal atoms at the cathode and nonmetal anions are oxidized at the anode.
- Electrorefining is a process used to obtain high-grade metals at the cathode from an impure metal at the anode.
- Electroplating is a process in which a metal is deposited on the surface of an object placed at the cathode of an electrolytic cell.

Section 14.3 Questions

1. Define an electrolytic cell.
2. List the key similarities and differences in the typical laboratory construction of a voltaic cell and an electrolytic cell.
3. The rules for predicting the chemical reactions, and ion and electron movements are essentially the same for both voltaic and electrolytic cells. What is different about the cell reaction equations and characteristics of each type of cell?
4. Two tin electrodes are placed in an aqueous solution containing potassium nitrate and magnesium iodide.
   (a) If a power supply is connected to force any reactions to occur, what would be the reactions at the cathode, anode, and in the overall cell?
   (b) Draw and label a cell diagram, including electrodes, electrolyte, power supply, and the direction of movement of electrons and ions.
   (c) Would a 1.5 V cell be suitable as a power supply? Justify your answer.
5. For the electrolysis of aqueous solutions, describe the common exception to the rules for predicting products of an electrolytic cell. Clearly identify the circumstances when this exception is used.
6. List three uses of electrolytic cells in industry.
7. Why were many metals discovered only after the invention of the electric cell?
8. How does the occurrence of metals in nature relate to the redox table?
9. Draw and label a simple cell for the electrolysis of molten potassium iodide (m.p. 682 °C). Label electrodes and power supply, directions of electron and ion flow, and write half-reaction and net equations.
10. When refining metals in an electrolytic cell, why must the metal product form at the cathode?
11. High-purity copper metal is produced using electrorefining.
    (a) At which electrode is the impure copper placed? Why?
    (b) What is the minimum electric potential difference required for this cell?
    (c) Why is it unlikely that your answer to (b) would be used? Discuss briefly.
12. How can you predict which metals might be refined from an aqueous solution?
13. Describe the relationship of science and technology in the area of electrolysis. Include several examples in your description.
14. "German silver" is an alloy containing copper, zinc, and nickel. A piece of German silver is used as the anode in an electrolytic cell containing aqueous sodium sulfate. The other electrode is platinum metal.
    (a) As the applied voltage is slowly increased, in what order will the half-reactions occur at the anode? Write an equation for each half-reaction.
    (b) Describe what happens at the cathode.
    (c) German silver does not contain any silver metal. Why is it called German silver? Why is it a very useful alloy?

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15. Design a cell to electroplate zinc onto an iron spoon. In your cell diagram, include:
    - ions in the solution
    - substances used for the electrodes
    - anode and cathode labels
    - power supply, showing signs and connections
    - directions of ion and electron flow
16. Suppose you work for a mining company and you are given a job to design a process that will recover nickel metal from a waste solution containing nickel(II) ions.
    (a) Propose an experimental design involving electrolysis that could be tested in the laboratory on a small scale.
    (b) What are some possible complications or factors that need to be considered? List these as questions.

Extension

17. The one-dollar coin, or the loonie (**Figure 13**) replaced the one-dollar bill, which typically wore out in a few months. Sherritt Gordon of Fort Saskatchewan, Alberta developed a unique process for plating the loonie coin.
    (a) Research the production and composition of the loonie.
    (b) What is the golden "aureate" finish on the loonie? Describe the materials and process for producing this finish.
    (c) Why did the coin end up with a loon stamped on it?

Figure 13
The Royal Canadian Mint in Winnipeg produces the Canadian loonie for general circulation.

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18. Aluminium cans are widely used to contain beverages. Write a short report about the production of aluminium cans, including how the can is made, how the top is attached to the can, how the construction of the can has changed since the first model, and the advantages of using recycled aluminium instead of new aluminium.
19. "Cold fusion" made front-page news when it was announced in 1989 as a new practical source of energy.
    (a) Briefly describe cold fusion.
    (b) How is cold fusion related to electrochemistry?
    (c) Outline some theoretical and empirical arguments in the controversy about the existence of cold fusion.

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14.4 Cell Stoichiometry

In the production of elements, the refining of metals, and electroplating, the quantity of electricity that passes through a cell determines the masses of substances that react or are produced at the electrodes. As you know from oxidation and reduction half-reactions, a specific number of electrons are lost or gained. For example, when zinc is plated onto a steel pipe to galvanize it, two moles of electrons must be gained by one mole of zinc ions to deposit one mole of zinc atoms as metal.

$$Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)$$

As in all stoichiometry, this relationship establishes a mole ratio of electrons to some other substance in the half-reaction equation. Unfortunately, there is no meter or instrument for measuring directly (or counting directly) the number of electrons. The number of electrons (as moles of electrons) is determined indirectly. In the past, you have measured mass and then converted to a chemical amount of a substance; a similar procedure is necessary for determining the amount of electrons.

Before we can look at the amount of electrons, we need to see how the charge is determined. Charge, Q, in coulombs, is determined from the electric current, I, in amperes (coulombs per second), and the time, t, in seconds, according to the following definition:

$$Q = It$$

One coulomb (C) is the quantity of charge transferred by a current of one ampere (A) during a time of one second

DID YOU KNOW ?

Science versus Technology

Jack Kilby (1923–2005) was a junior engineer when he invented the microchip that is the foundation of all modern electronics. He was awarded a Nobel Prize for Physics in 2000. On the topic of his prize, Kilby commented:

"Those big prizes are for the advancement of understanding. They are for scientists who are motivated by pure knowledge. But I'm an engineer. I'm motivated by a need to solve problems, to make something work. For guys like me, the prize is seeing a successful solution."

SAMPLE problem 14.2

The technology of the Hall–Héroult cell for producing aluminium has improved considerably since the first industrial factory. Modern electrolytic cells may use up to 300 kA of current. What is the charge that passes through one of these cells in a 24 h period?

By definition, a current in amperes (A) is the number of coulombs per second, 1 A = 1 C/s. You always need to convert the time into seconds before time can be used in the calculation of charge.

$$t = 24\ \cancel{h} \times \frac{3600\ s}{1\ \cancel{h}} = 8.6 \times 10^4\ s$$

Now the charge in coulombs can be calculated as follows:

$$\begin{aligned} Q &= It \\ &= 300 \times 10^3\ \frac{C}{\cancel{s}} \times 8.6 \times 10^4\ \cancel{s} \\ &= 2.6 \times 10^{10}\ C \end{aligned}$$

Therefore, a current of 300 kA for 24 h transfers 2.6×10^{10} C of charge. This is a huge quantity of charge. For comparison, the charge passing through a 100 W light bulb in 24 h is about 7.2×10^4 C.

Practice

1. Calculate the charge transferred by a current of 1.5 A flowing for 30 s.
2. In an electrolytic cell, 87.6 C of charge is transferred in 22.5 s. Determine the electric current.
3. Calculate the charge transferred by a current of 250 mA in a time of 28.5 s.
4. How long, in minutes, does it take a current of 1.60 A to transfer a charge of 375 C?

Faraday's Law

The relationship between electricity and electrochemical changes was first investigated by Michael Faraday (**Figure 1**) in the 1830s. Faraday continued Humphry Davy's work in electrochemistry, coining the terms electrolysis, electrolyte, electrode, anode, cathode, cation, and anion. His quantitative study of electrolysis identified the factors that determine the mass of an element produced or consumed at an electrode. He discovered that this mass was directly proportional to the time the cell operated, as long as the current was constant (**Faraday's law**). Furthermore, he found that *9.65 × 10^4 C of charge is transferred for every mole of electrons that flows in the cell.* In modern terms, this value is the molar charge of electrons, also called the **Faraday constant**, *F*.

$$F = 9.65 \times 10^4 \frac{\text{C}}{\text{mol e}^-}$$

This constant can be used as a conversion factor in converting electric charge to an amount in moles of electrons—in the same way that molar mass is used to convert mass to a chemical amount.

$$n_{e^-} = \frac{Q}{F}$$

Since $Q = It$, the amount of electrons can now be written as

$$n_{e^-} = \frac{It}{F}$$

Figure 1
As a young man, Michael Faraday (1791–1867) taught himself chemistry and convinced the famous English chemist Humphry Davy to hire him as his assistant. Faraday proved himself more than worthy of this trust. He eventually made important contributions in the study of gases, low temperatures, the discovery of benzene, quantitative aspects of electrolysis, electric motors, generators, and transformers. A deeply religious man, Faraday had strong convictions about the appropriate uses of science and technology. He refused to help Britain produce a poison gas for use against the Russians in the Crimean War (1854–56).

SAMPLE problem 14.3

What amount of electrons is transferred in a cell that operates for 1.25 h at a current of 0.150 A?

Recall from the calculation of electric charge that the time must always be in seconds because the ampere is defined as coulombs per second (1 A = 1 C/s).

$$t = 1.25 \cancel{\text{h}} \times \frac{3600 \text{ s}}{1 \cancel{\text{h}}} = 4.50 \times 10^3 \text{ s}$$

Now you can calculate the amount, in moles, of electrons using the Faraday constant as a conversion factor:

$$n_{e^-} = (0.150 \frac{\cancel{\text{C}}}{\cancel{\text{s}}} \times 4.50 \times 10^3 \cancel{\text{s}}) \times \frac{1 \text{ mol}}{9.65 \times 10^4 \cancel{\text{C}}}$$

$$= 6.99 \times 10^{-3} \text{ mol}$$

The amount of electrons transferred is 6.99 mmol.

Learning Tip

Note the similarity between various definitions of amounts using molar quantities:

$$n_{\text{substance}} = \frac{m}{M}$$

$$n_{\text{gas}} = \frac{v}{V}$$

$$n_{e^-} = \frac{Q}{F}$$

COMMUNICATION example 1

Convert a current of 1.74 A for 10.0 min into an amount of electrons.

Solution

$$t = 10.0 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} = 600 \text{ s}$$

$$n_{e^-} = (1.74 \frac{\text{C}}{\text{s}} \times 600 \text{ s}) \times \frac{1 \text{ mol}}{9.65 \times 10^4 \text{ C}}$$

$$= 0.0108 \text{ mol}$$

According to Faraday's law, the amount of electrons transferred is 0.0108 mol or 10.8 mmol.

The same method can be used to calculate electric current or time if the other variables are known, as shown in Communication Example 2.

Learning Tip

An alternative solution is to use the mathematical formula and solve for time, t.

$$n_{e^-} = \frac{It}{F}$$

$$t = \frac{n_{e^-}F}{I}$$

You can either substitute values first and then solve for time, or rearrange the mathematical formula and then substitute the values. Either way is acceptable.

COMMUNICATION example 2

How long, in minutes, will it take a current of 3.50 A to transfer 0.100 mol of electrons?

Solution

$$t = (0.100 \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}) \times \frac{1 \text{ s}}{3.50 \text{ C}}$$

$$= 2.76 \times 10^3 \text{ s} \times \frac{1 \text{ min}}{60 \text{ s}}$$

$$= 46.0 \text{ min}$$

According to Faraday's law, it would require 46.0 min to transfer 0.100 mol e^- using a current of 3.50 A.

Practice

5. An electroplating cell operates for 35 min with a current of 1.9 A. Calculate the amount, in moles, of electrons transferred.
6. A cell transferred 0.146 mol of electrons with a constant current of 1.24 A. How long, in hours, did this take?
7. Calculate the current required to transfer 0.015 mol of electrons in 20 min.

Half-Cell Calculations

Since the mass of an element produced at an electrode depends on the amount of transferred electrons, a half-reaction equation showing the number of electrons involved is necessary to do stoichiometric calculations. This applies to all electrochemical cells, whether voltaic or electrolytic. Separate calculations are carried out for each electrode, although the *same charge and, therefore, the same amount of electrons passes through each electrode in a cell* or a group of cells in series. As the following examples show, concepts of stoichiometry used in other calculations also apply to half-cell calculations. The only new part of the stoichiometry is the calculation of the amount of electrons based on the Faraday constant.

SAMPLE problem 14.4

What is the mass of copper deposited at the cathode of a copper electrorefining cell (**Figure 2**) operated at 12.0 A for 40.0 min?

First identify and write the appropriate half-cell equation. Because copper is being deposited at the cathode, copper(II) ions must be gaining electrons to form copper metal. Write the equation for this reduction and list all information given, including constants such as molar mass and Faraday.

$$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$$

40.0 min — *m*
12.0 A — 63.55 g/mol
9.65×10^4 C/mol

Notice that we have all of the information necessary to calculate the amount of electrons. Don't forget to make sure the time is converted to units of seconds, if necessary.

$$n_{e^-} = \left(12.0 \frac{\cancel{C}}{s} \times 40.0 \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}}\right) \times \frac{1\ mol}{9.65 \times 10^4\ \cancel{C}}$$

$$= 0.298\ mol$$

The procedure that is common to all stoichiometry is the use of the mole ratio from a balanced equation. The mole ratio is what allows us to convert from a chemical amount in moles of one substance to another. In the reduction half-reaction given, notice that 1 mol of copper metal is formed when 2 mol of electrons are transferred.

$$n_{Cu} = 0.298\ mol \times \frac{1}{2}$$

$$= 0.149\ mol$$

The final step is to convert to the quantity requested in the question, in this case, the mass of copper metal.

$$m_{Cu} = 0.149\ \cancel{mol} \times \frac{63.55\ g}{\cancel{mol}}$$

$$= 9.48\ g$$

or

$$m_{Cu} = 40.0\ \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times 12.0 \frac{\cancel{C}\ e^-}{1\ \cancel{s}} \times \frac{1\ mol\ e^-}{9.65 \times 10^4\ \cancel{C}\ e^-} \times \frac{1\ mol\ Cu}{2\ mol\ e^-} \times \frac{63.55\ g\ Cu}{1\ mol\ Cu}$$

$$= 9.48\ g$$

According to Faraday's law and the stoichiometric method, the mass of copper metal deposited is 9.48 g.

Figure 2
These pure copper cathodes from the electrolytic refining cells are ready for melting and processing into copper wire and other copper products.

Learning Tip

Note the similarity of the procedure for stoichiometry calculations of half-cells to all other stoichiometry calculations you have done in the past. Essentially, the only difference is a new relationship (formula based on Faraday's law) to convert to and from the amount of electrons.

SUMMARY Procedure for Half-Cell Stoichiometry

Step 1: Write the balanced equation for the half-cell reaction of the substance produced or consumed. List the measurements and conversion factors for the given and required entities.

Step 2: Convert the given measurements to an amount in moles by using the appropriate conversion factor (M, c, F).

Step 3: Calculate the amount of the required substance by using the mole ratio from the half-reaction equation.

Step 4: Convert the calculated amount to the final quantity by using the appropriate conversion factor (M, c, F).

Figure 3
A silver electroplating cell uses silver cyanide to silver plate objects, such as this tray.

COMMUNICATION example 3

Silver is deposited on objects (**Figure 3**) in a silver electroplating cell. If 0.175 g of silver is to be deposited from a silver cyanide solution in a time of 10.0 min, predict the current required.

Solution

$$\begin{array}{ccccc} Ag^+(aq) & + & e^- & \rightarrow & Ag(s) \\ & & 10.0\ min & & 0.175\ g \\ & & I & & 107.87\ g/mol \\ & & 9.65 \times 10^4\ C/mol & & \end{array}$$

$$n_{Ag} = 0.175\ \cancel{g} \times \frac{1\ mol}{107.87\ \cancel{g}}$$

$$= 1.62 \times 10^{-3}\ mol$$

$$n_{e^-} = 1.62 \times 10^{-3}\ mol \times \frac{1}{1}$$

$$= 1.62 \times 10^{-3}\ mol$$

$$I = \frac{1.62 \times 10^{-3}\ \cancel{mol} \times 9.65 \times 10^4\ \frac{C}{\cancel{mol}}}{10.0\ \cancel{min} \times \frac{60\ s}{\cancel{min}}}$$

$$= 0.261\ C/s$$

or

$$I = 0.175\ \cancel{g\ Ag} \times \frac{1\ \cancel{mol\ Ag}}{107.87\ \cancel{g\ Ag}} \times \frac{1\ \cancel{mol\ e^-}}{1\ \cancel{mol\ Ag}} \times \frac{9.65 \times 10^4\ C\ e^-}{1\ \cancel{mol\ e^-}} \times \frac{1}{10\ \cancel{min}} \times \frac{1\ \cancel{min}}{60\ s}$$

$$= 0.261\ C/s$$

According to the stoichiometry and Faraday's law, the current required to plate 0.175 g of silver in 10.0 min is 0.261 A.

WEB Activity

Simulation—Electrolytic Cell Stoichiometry

In this activity, you will simulate the operation of an electrolytic cell. You will follow a series of steps to calculate the approximate cost of producing a given mass of aluminium. These calculations are similar to ones that scientists and engineers do when designing and monitoring an industrial cell.

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Section 14.4 Questions

1. A battery delivers 0.300 A for 15.0 min. What amount of electrons, in moles, is transferred?
2. A current of 55 kA passes through a chlor–alkali cell. What mass of chlorine is formed during 8.0 h ?
3. A family wishes to plate an antique teapot with 10.00 g of silver. If the current to be used is 1.80 A, what length of time, in minutes, is required?
4. A typical Hall–Héroult cell produces 425 kg of molten aluminium in 24.0 h. Calculate the current used.
5. Magnesium metal is produced in an electrolytic cell containing molten magnesium chloride. A current of 2.0×10^5 A is passed through the cell for 18.0 h.
   (a) Determine the mass of magnesium produced.
   (b) What mass of chlorine is produced at the same time?
6. Cobalt metal is plated from 250.0 mL of cobalt(II) sulfate solution. What is the minimum concentration of cobalt(II) sulfate required for this cell to operate for 2.05 h with a current of 1.14 A?
7. A 25.72 g piece of copper metal is the anode in a cell in which a current of 0.876 A flows for 75.0 min. Determine the final mass of the copper electrode.
8. A student reconstructs Volta's electric battery using sheets of copper and zinc, and a current of 0.500 A is produced for 10.0 min. Calculate the mass of zinc oxidized to aqueous zinc ions.
9. Electroplating is a common technological process for coating objects with a metal to enhance the appearance of the object or its resistance to corrosion.
   (a) A car bumper is plated with chromium using chromium(III) ions in solution. If a current of 54 A flows in the cell for 45 min 30 s, determine the mass of chromium deposited on the bumper.
   (b) For corrosion resistance, a steel bolt is plated with nickel from a solution of nickel(II) sulfate. If 0.250 g of nickel produces a plating of the required thickness and a current of 0.540 A is used, predict how long in minutes the process will take.
10. During the electrolysis of molten aluminium chloride in an electrolytic cell, 5.40 g of aluminium is produced at the cathode. Predict the mass of chlorine produced at the anode.
11. Chromium metal can be plated onto an object from an acidic solution of dichromate ions. What average current is required to plate 17.8 g of chromium metal in a time of 2.20 h? (You will need to construct your own equation for the half-reaction.)
12. The purpose of this experiment is to test the method of stoichiometry in cells. Complete the Prediction, Analysis, and Evaluation (Part 2 only) of the investigation report.

    Problem
    What is the mass of tin electroplated at the cathode of a tin-plating cell by a current of 3.46 A for 6.00 min?

    Design
    A steel can is placed in an electroplating cell as the cathode. An electric current of 3.46 A flows through the cell, which contains a 3.25 mol/L solution of tin(II) chloride, for 6.00 min.

    Evidence
    initial mass of can = 117.34 g
    final mass of can = 118.05 g
13. Using a specific example of an electrolytic cell, describe how Faraday's law is useful in designing and controlling the process.

Extension

14. A rapidly developing technology is the production of less expensive, more durable, and more energy-dense electrochemical cells, that is, cells with a high energy-to-mass ratio.
    (a) A car battery has a rating of 125 A•h (ampere-hours). What does this tell you about the electrical capacity of this battery?
    (b) Why is this a useful way to rate batteries?
    (c) What mass of lead is oxidized as this battery discharges?
    (d) If an aluminium–oxygen fuel cell has the same rating as the car battery in (a), what mass of aluminium metal would be oxidized?
    (e) Comment on the implications of your answers to (c) and (d).

INVESTIGATION 14.1

Designing an Electric Cell

In this cell, an aluminium soft-drink can is one of the electrodes (**Figure 1**). The other electrode is a solid conductor, such as a piece of copper wire or pipe, an iron nail, or graphite from a pencil. The electrolyte may be a salt solution, or an acidic or basic solution. Although many characteristics of a cell are important, only one characteristic, voltage, is investigated here. Check with your teacher if you want to evaluate other designs and materials.

When evaluating the Purpose (Part 3), include your opinion about the reliability, cost, and simplicity of your final electric cell.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	● Materials	● Evaluation (1, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Be careful when handling acidic and basic solutions used for electrolytes, as they are corrosive. Wear eye protection and work near a source of water. Some electrolytes may be toxic or irritant; follow all safety precautions. Avoid eye and skin contact.

Purpose

The purpose of this investigation is to make an electric cell.

Problem

What combination of electrodes and electrolyte gives the largest voltage for an aluminium-can cell?

Design

Part 1

Using the same electrolyte and aluminium can as the controlled variables, two or three different materials are used as the second electrode. The voltage of each cell is measured as the responding variable.

Part 2

Using the same two electrodes as the controlled variables, two or three possible electrolytes are tested. The voltage of each cell is measured as the responding variable.

Part 3

Additional combinations are tested, based on the analysis of the initial trials.

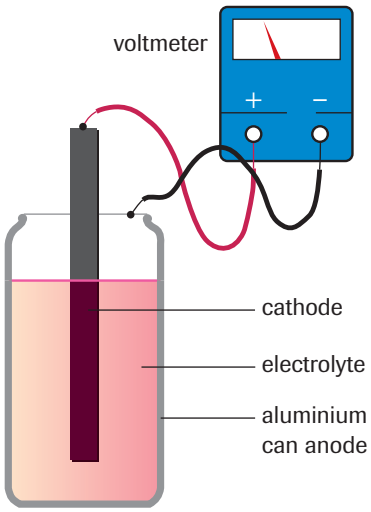


Figure 1
An aluminium-can cell is an efficient design since one of the electrodes also serves as the container.

INVESTIGATION 14.2

A Voltaic Cell (Demonstration)

An important characteristic of consumer, commercial, and industrial cells is a simple, efficient design that works for the intended application. For scientific research, this is not as important as a design that can be easily manipulated and studied.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the design and operation of a voltaic cell used in scientific research.

Problem

What is the design and operation of a voltaic cell?

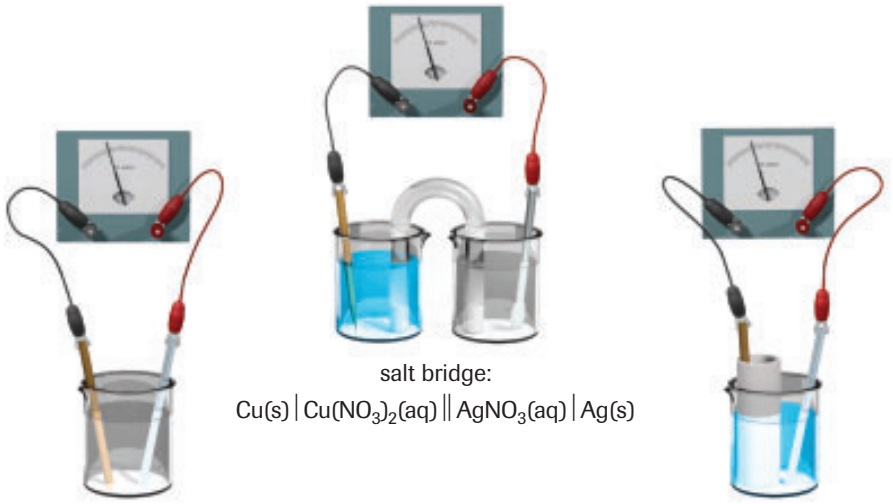
Design

An electric cell with only one electrolyte is compared with voltaic cells containing the same electrodes, but two electrolytes.

Solutions used are toxic and irritant.
Avoid contact with skin and eyes.

Procedure

1. Construct the three cells shown in **Figure 2**.
2. For each design, use a voltmeter to determine which electrode is positive and which is negative (see Appendix C.3), and measure the electric potential difference of each cell.
3. With the voltmeter connected, remove and then replace the various parts of the cell.
4. For each cell, connect the two electrodes with a wire. Record any evidence of a reaction after several minutes, and after one or two days. Measure the electric potential difference after several days.



no porous boundary:
$Cu(s) \mid NaNO_3(aq) \mid Ag(s)$

salt bridge:
$Cu(s) \mid Cu(NO_3)_2(aq) \parallel AgNO_3(aq) \mid Ag(s)$

porous cup:
$Cu(s) \mid Cu(NO_3)_2(aq) \parallel AgNO_3(aq) \mid Ag(s)$

Figure 2
In Investigation 14.2 , you compare three different cell designs.

INVESTIGATION 14.3

Testing Voltaic Cells

Testing is a procedure that is common to both technology and science. In technology, testing is necessary to determine how a product or process works using criteria such as efficiency, reliability, and cost. In science, testing is a key part in the advancement of knowledge. Scientific concepts are developed and then tested to determine their validity and limitations. New ideas that fail the test then need to be restricted, revised, or replaced.

In your Evaluation, pay particular attention to sources of error or uncertainty, and to limitations of the evidence collected.

Purpose

The purpose of this investigation is to test the predictions of cell potentials and the identity of the electrodes of various cells.

Problem

In cells constructed from various combinations of copper, aluminium, silver, and zinc half-cells, what are the standard cell potentials, and which is the anode and cathode in each case?

The materials used are toxic and irritant. Avoid skin and eye contact.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Materials

lab apron
eye protection
voltmeter and connecting wires
U-tube with cotton plugs, porous cups, or filter paper
four 100 mL beakers (or well plate)
distilled water
steel wool
Cu(s), Al(s), Ag(s), and Zn(s) strips
1.0 mol/L each of:
- $CuSO_4(aq)$
- $Al(NO_3)_3(aq)$
- $AgNO_3(aq)$
- $NaNO_3(aq)$
- $ZnSO_4(aq)$

INVESTIGATION 14.4

A Potassium Iodide Electrolytic Cell

Electrolytic cells were discovered before the science was understood. However, as with all successful technological inventions, the important criteria was that it worked, not why it worked. Eventually, chemists understood the science and were able to explain why electrolytic cells work.

In the Evaluation, suggest changes to the Design, Materials, and Procedure that would improve the Evidence.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 3)
○ Hypothesis	○ Procedure	
○ Prediction	● Evidence	

Purpose

The purpose of this investigation is to use diagnostic tests to determine the reaction products of an electrolytic cell.

Problem

What are the products of the reaction during the operation of an aqueous potassium iodide electrolytic cell?

Design

Inert electrodes are placed in a 0.50 mol/L solution of potassium iodide, and a battery or power supply provides a direct current of electricity to the cell. The litmus and halogen diagnostic tests (Appendix C.4) are conducted to test the solution near each electrode before and after the reaction.

Materials

lab apron
eye protection
petri dish
two carbon electrodes
two connecting wires
3 V to 9 V battery or power supply
red and blue litmus paper
ring stand and two utility clamps
small test tube with stopper
dropper bottle of hexane
0.50 mol/L KI(aq)

Hexane is highly flammable. Do not use near an open flame and avoid inhaling the fumes. Use only in a well-ventilated area.

Procedure

1. Set up the KI(aq) cell as shown in **Figure 3**, (or with one ring stand), but with a single wire connecting the electrodes (i.e., no power supply).
2. Observe the cell and test the solution with litmus paper and hexane.
3. Use two connecting wires to hook up the power supply to the electrodes.
4. Turn on the power supply.
5. Record all observations at each electrode.
6. Perform both diagnostic tests at each electrode.
7. Dispose of the solutions by putting the hexane mixture in a labelled waste container and washing the potassium iodide solution down the sink drain.

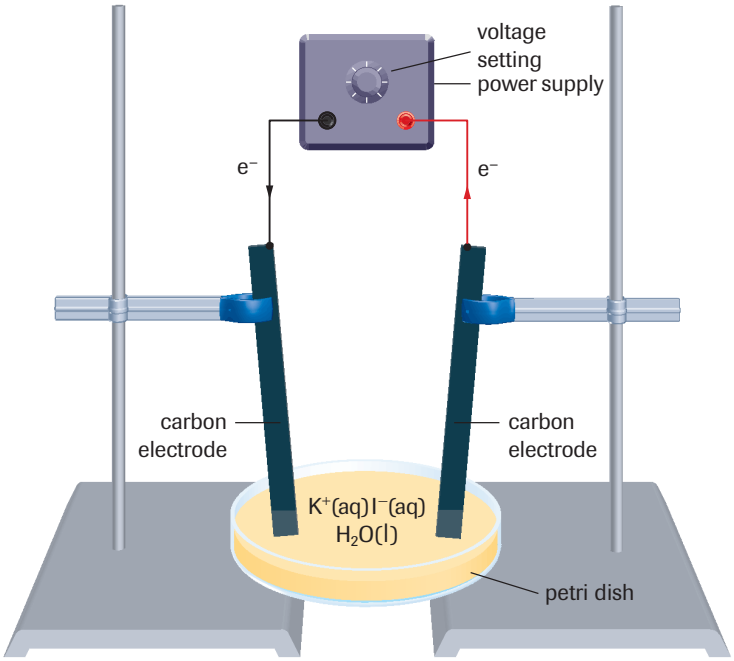


Figure 3
Apparatus for the electrolysis of potassium iodide

INVESTIGATION 14.5

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ● Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

Electrolysis (Demonstration)

Scientific knowledge progresses by the experimental testing of ideas. The more rigorous the test, the more certain the knowledge or the better the chance of making new discoveries.

Purpose

The purpose of this demonstration is to test the method of predicting the products of electrolytic cells.

Problems

What are the products of electrolytic cells containing one of the following solutions:

- aqueous copper(II) sulfate?
- aqueous sodium sulfate?
- aqueous sodium chloride?

Design

The electrolysis of the aqueous copper(II) sulfate is carried out in a U-tube, and the electrolysis of aqueous sodium sulfate and sodium chloride is carried out in a Hoffman apparatus (**Figure 4**) so that any gases produced can be collected. Diagnostic tests with necessary control tests (before electrolysis) are conducted to determine the presence of the predicted products.

Copper(II) sulfate is toxic and irritant. Avoid skin and eye contact. If you spill copper(II) sulfate solution on your skin, wash the affected area with lots of cool water. During electrolysis, corrosive substances are produced; avoid skin and eye contact.

Figure 4
Hoffman apparatus

Chapter 14 SUMMARY

Outcomes

Knowledge

- define anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, voltaic cell, and electrolytic cell (14.2, 14.3)
- predict and write the half-reaction equation that occurs at each electrode in an electrochemical cell (14.2, 14.3, 14.4)
- explain that the values of standard reduction potential are all relative to $E°_r = 0.00$ V set for the hydrogen electrode at standard conditions (14.2)
- calculate the standard cell potential for electrochemical cells (14.2, 14.3)
- predict the spontaneity of redox reactions based on standard cell potentials (14.2, 14.3)
- identify the similarities and differences between a voltaic cell and an electrolytic cell (14.3)
- recognize that predicted reactions do not always occur (14.3)
- calculate mass, amounts, current, and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry (14.4)

STS

- state examples of science leading technology and technology leading science (14.1, 14.3)
- recognize the values and limitations of technological products and processes (14.1, 14.3)
- describe the interactions of science and technology (14.1, 14.2, 14.3)

Skills

- initiating and planning: design an experiment, including a labelled diagram, to test predictions for reactions occurring in electrochemical cells (14.2, 14.3); describe procedures for safe handling, storage, and disposal of materials used in the laboratory (14.1, 14.2); and develop a plan to build a cell (battery) using a trial-and-error procedure (14.1)
- performing and recording: construct and observe electrochemical cells (14.1, 14.2, 14.3); investigate the issue of disposal of batteries (14.1); and compile and display information about electrochemical cells (all sections)
- analyzing and interpreting: identify the products of electrochemical cells (all sections); compare predictions with observations (all sections); identify the limitations of evidence collected (14.1, 14.2, 14.3); explain discrepancies between predicted and measured cell potentials (14.2, 14.3); assess the practicalities of different cell designs (14.1); and evaluate experimental designs for cells (14.1, 14.2, 14.3)
- communication and teamwork: work collaboratively in addressing problems and use appropriate SI notation and significant digits (all sections); and select and integrate information from various sources about technological applications of cells (all sections)

Key Terms

14.1

electrode
electrolyte
electric potential difference
electric current
fuel cell
standard cell
standard cell potential
standard reduction potential
reference half-cell
corrosion
cathodic protection

14.2

porous boundary
half-cell
voltaic cell
cathode
anode
inert electrode

14.3

electrolytic cell
electrolysis

14.4

Faraday's law
Faraday constant

Key Equations

- $E°_{cell} = E°_{r\ cathode} - E°_{r\ anode}$ (14.2, 14.3)
- $n_{e^-} = \frac{It}{F}$ (14.4)

MAKE a summary

1. Draw a diagram of a simple, general voltaic cell. Show and label all components, including names and signs where appropriate. Show the directions of electron and ion movement. State the process occurring at each electrode and how the cell potential is determined. List several specific technological examples.
2. Repeat question 1 for a simple, general electrolytic cell.
3. Refer back to your answers to the Starting Points questions at the beginning of this chapter. How has your thinking changed?

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The following components are available on the Nelson Web site. Follow the links for *Nelson Chemistry Alberta 20–30*.

- an interactive Self Quiz for Chapter 14
- additional Diploma Exam-style Review questions
- Illustrated Glossary
- additional IB-related material

There is more information on the Web site wherever you see the Go icon in this chapter.

Chapter 14 REVIEW

Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

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DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. Which one of the following is the best example of a new technology that led to scientific discoveries?
A. electric cell
B. Faraday constant
C. reference half-cell
D. hydrogen fuel cell

2. Which of the following statement applies more to science than to technology?
A. It involves the development of devices/process that have a practical purpose.
B. It seeks explanations about the natural world.
C. It applies mainly to the human-designed processes.
D. It applies trial-and-error strategies.

3. The main advantage of fuel cells is that they
A. are reliable
B. are portable
C. are inexpensive to build
D. run continuously as reactants are added

4. In a voltaic cell, the cathode is labelled as the __*i*__ electrode where the __*ii*__ half reaction of the strongest __*iii*__ agent occurs.

The above statement is completed by the information in which row?

Row	*i*	*ii*	*iii*
A.	negative	oxidation	oxidizing
B.	negative	oxidation	reducing
C.	positive	reduction	oxidizing
D.	positive	reduction	reducing

Use the following information to answer questions 5 and 6.
1. net reaction is spontaneous
2. net reaction is nonspontaneous
3. anions migrate toward the anode
4. cell potential is positive
5. cell potential is negative
6. electrons flow from the cathode to the anode
7. ions complete the internal circuit in the cell

5. NR In numerical order, the statements that are true for voltaic cells are ___ ___ ___ ___.

6. NR In numerical order, the statements that are true for electrolytic cells are ___ ___ ___ ___.

7. NR The cell potential of a standard lead–chlorine cell is predicted to be

+_____ V

8. The main products formed in the electrolysis of an aqueous nickel(II) chloride solution are

	at the anode	at the cathode
A.	$O_2(g)$, $H^+(aq)$	Ni(s)
B.	$Cl_2(g)$	Ni(s)
C.	$Cl_2(g)$	$H_2(g)$, $OH^-(aq)$
D.	$Ni^{2+}(aq)$	$Cl^-(aq)$

9. If chemists had decided to use the standard zinc half-cell as the reference half-cell, the new standard reduction potential for $Hg^{2+}(aq)$ would be
A. −1.61 V
B. −0.09 V
C. +0.09 V
D. +1.61 V

10. What is the standard reduction potential for $Ce^{4+}(aq)$ using the following net reaction and standard cell potential?

$$Ce^{4+}(aq) + Fe^{2+}(aq) \rightarrow Ce^{3+}(aq) + Fe^{3+}(aq) \qquad E^\circ_{cell} = +0.67\ V$$

A. +1.44 V
B. +0.10 V
C. −0.10 V
D. −1.44 V

11. The industrial production of alkali metals involves
A. electrolysis of molten salts
B. electrolysis of aqueous salts
C. reduction with strong reducing agents
D. oxidation with strong oxidizing agents

12. Which of the following is **not** an effective method of corrosion prevention?
A. galvanizing
B. sacrificial anode
C. sacrificial cathode
D. impressed current

13. NR In an electroplating operation, the mass of copper metal produced from a copper(II) ion solution by a 0.325 A current flowing for 3.00 h is predicted to be

_____ g.

14. NR The following electrolytic cell is set up:

$$C(s) \mid Co(NO_3)_2(aq) \mid C(s)$$

The minimum voltage that must be applied to operate this cell is

+_____ V

Part 2

15. **Explain** the function of the following components of voltaic cells.
(a) electrolyte
(b) salt bridge
(c) electrode
(d) connecting wire

16. For each of the following cells, write the equations for the reactions occurring at the cathode and at the anode, and the equation for the net cell reaction.
(a) $Ni(s) \mid Ni^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$
(b) $Zn(s) \mid Zn^{2+}(aq) \parallel Cr_2O_7^{2-}(aq), H^+(aq) \mid C(s)$
(c) $Pt(s) \mid H_2(g) \mid OH^-(aq) \parallel Ag^+(aq) \mid Ag(s)$

17. Draw and label a diagram of a voltaic cell constructed from some (not all) of the following materials:

strip of silver metal	voltmeter
strip of lead metal	connecting wires
aqueous silver nitrate	glass U-tube
aqueous lead(II) nitrate	cotton
aqueous sodium nitrate	various beakers
distilled water	porous porcelain cup

18. Provide the following information for the cell constructed in question 17.
(a) anode half-reaction
(b) cathode half-reaction
(c) net cell reaction
(d) standard cell potential

19. DE Cathodic protection is widely used in industry for preventing the corrosion of steel tanks and pipes.
(a) **Describe** two forms of cathodic protection including the chemical principles involved.
(b) Assess the importance of corrosion protection using at least two perspectives.

20. A student constructs an electrolytic cell to nickel plate a carbon rod. Draw and label a diagram of an electrolytic cell constructed using the following materials:

strip of nickel metal	power source
carbon rod	connecting wires
aqueous nickel(II) sulfate	large beaker
voltmeter	

21. Provide the following information for the cell constructed in question 20.
(a) anode half-reaction
(b) cathode half-reaction
(c) net cell reaction
(d) minimum voltage required

22. DE As part of a class project, a student needs to construct a cell that produces a voltage sufficient to operate a 1.25 V clock.
(a) **Outline** the design of a suitable cell including a fully labelled cell diagram.
(b) **Justify** your design using appropriate chemical terms, equations, and calculations.
(c) Suggest some possible discrepancies or limitations of this cell for the chosen task.

23. In the Hall–Héroult process, an electric current is passed through the electrolyte at low voltage, but very high current. Calculate the mass of aluminium metal that would be produced from aluminium ions by a 150 kA current flowing in a Hall–Héroult cell for 1.00 h (**Figure 1**).

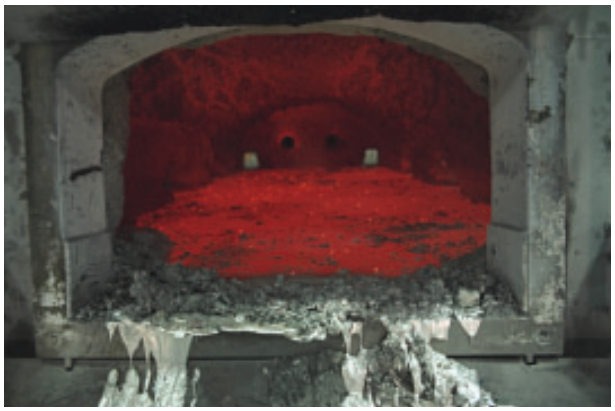

Figure 1
Molten aluminium from an electrolytic cell

24. A teacher constructs a demonstration voltaic cell, $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$, and allows it to operate for 24.0 h. The mass of the cathode was measured before and after the demonstration. Use the evidence collected to calculate the average current produced by this cell.

final mass of the cathode:	24.68 g
initial mass of the cathode:	21.12 g

25. DE Briefly **explain** the process of corrosion. Your response should include
- several examples of metal corrosion of manufactured materials
- any environmental, health, or safety issues associated with your examples
- brief descriptions of some technological solutions to the problem of corrosion
- at least one example of corrosion that is desirable

Unit 7 REVIEW

Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

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DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. A redox reaction involves a transfer of electrons
A. from the oxidizing agent to the reducing agent
B. from the reducing agent to the oxidizing agent
C. through a porous barrier
D. between metals only

2. A general reaction type that is **not** a redox reaction is
A. neutralization
B. disproportionation
C. combustion
D. formation

3. When solutions
NR
1. sulfuric acid
2. lithium hydroxide
3. gold(III) fluoride
4. chromium(II) nitrate

are ranked in order of strength of oxidizing agents, the order, from **strongest to weakest** oxidizing agent, is __, __, __, and __.

4. During the process of photosynthesis,

$6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$

A. carbon in carbon dioxide is oxidized
B. hydrogen in water is reduced
C. oxygen in carbon dioxide and/or water is oxidized
D. oxygen in glucose is oxidized

5. Which of the following reaction equations describes a redox reaction?
A. $C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$
B. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
C. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
D. $HMnO_4(aq) \rightarrow H^+(aq) + MnO_4^-(aq)$

6. The metal molybdenum, Mo(s), reacts to form $MoO_2(s)$. The half-reaction equation that explains the change in oxidation state of molybdenum can be written as
A. $Mo(s) + 2\ e^- \rightarrow Mo^{2+}(s)$
B. $Mo(s) \rightarrow Mo^{2+}(s) + 2\ e^-$
C. $Mo^{4+}(s) + 4\ e^- \rightarrow Mo(s)$
D. $Mo(s) \rightarrow Mo^{4+}(s) + 4\ e^-$

7. A high school laboratory's waste container is used to dispose of aqueous solutions of sodium nitrate, potassium sulfate, hydrochloric acid, and tin(II) chloride. The most likely net redox reaction predicted to occur inside the waste container is represented by the equation:
A. $2\ H^+(aq) + 2\ K^+(aq) \rightarrow H_2(g) + K(s)$
B. $Sn^{2+}(aq) + 2\ NO_3^-(aq) + 4\ H^+(aq) \rightarrow N_2O_4(g) + 2\ H_2O(l) + Sn^{4+}(aq)$
C. $SO_4^{2-}(aq) + 4\ H^+(aq) + 2\ Cl^-(aq) \rightarrow H_2SO_3(aq) + H_2O(l) + Cl_2(g)$
D. $2\ Cl^-(aq) + Sn^{2+}(aq) \rightarrow 2\ Cl_2(g) + Sn(s)$

Use this information to answer questions 8 to 11.

Steel is the most widely used alloy in the world. However, an estimated 20% of the iron and steel produced annually is used to replace that lost by corrosion through exposure to air and water. Therefore, corrosion prevention is of considerable importance. The empirical and theoretical chemistry of corrosion, and technological research and development together provide solutions to this important practical problem.

8. In the corrosion of steel objects in the natural environment, the most likely reducing agent is
A. $Fe^{2+}(aq)$
B. $Fe^{3+}(aq)$
C. Fe(s)
D. $O_2(g)$

9. The metals
NR
1. Fe(s)
2. Sn(s)
3. Zn(s)
4. Mg(s)

listed in order from *least to most* likely to corrode under similar atmospheric conditions are __, __, __, and __.

10. The reduction half-reaction that is generally involved in the corrosion of iron is
A. $Fe^{2+}(aq) + 2\ e^- \rightarrow Fe(s)$
B. $Fe(s) \rightarrow Fe^{3+}(aq) + 3\ e^-$
C. $2\ H_2O(l) \rightarrow O_2(g) + 4\ H^+(aq) + 4\ e^-$
D. $O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$

11. The net cell potential, under standard conditions, for the
NR
iron–oxygen cell in an aqueous environment is
+_____ V.

---

12. In a titration experiment, 10.0 mL samples of 0.650 mol/L
NR
chromium(II) ion solution reacted with an average volume of 12.4 mL of acidic potassium dichromate solution. The amount concentration of the potassium dichromate solution is
_____ mmol/L.

13. All voltaic and electrolytic cells require
A. one electrode and two electrolytes
B. two electrodes and one or two electrolytes
C. an external power supply
D. a voltmeter

14. Standard reduction potentials for half-cells are based on the strengths of
A. oxidizing agents relative to hydrogen ions
B. oxidizing agents relative to hydrogen gas
C. reducing agents relative to hydrogen ions
D. reducing agents relative to a standard acidic solution

15. In a voltaic cell, the reduction potentials of two standard half-cells are +0.35 V and −1.13 V. The predicted cell potential of the cell constructed from these two half-cells is
A. +0.35 V
B. +0.78 V
C. +1.13 V
D. +1.48 V

16. If the electrodes of a standard copper–silver cell are connected with a wire, then
A. silver is plated at the anode
B. a voltmeter would show a reading of 1.14 V
C. the solution at the anode becomes darker blue
D. electrons flow from the silver to the copper electrodes

17. The electrolysis of brine, NaCl(aq), is an important industrial process. The major products formed at each electrode are

	Cathode	Anode
A.	$H_2(g)$, $OH^-(aq)$	$O_2(g)$, $H^+(aq)$
B.	$H_2(g)$, $OH^-(aq)$	$Cl_2(g)$
C.	$Na(s)$	$O_2(g)$, $H^+(aq)$
D.	$Cl_2(g)$	$OH^-(aq)$

Use this information to answer questions 18 to 20.

An aqueous solution of potassium hydroxide undergoes electrolysis using 5.9 A of current for a total time of 22 min.

18. Electrons are transferred through the
A. solution from the anode to the cathode
B. solution from the cathode to the anode
C. external wire from the anode to the cathode
D. external wire from the cathode to the anode

19. The product(s) at the anode will be
A. $O_2(g)$, $H^+(aq)$
B. $K(s)$
C. $H_2(g)$, $OH^-(aq)$
D. $O_2(g)$, $H_2O(l)$

20. NR The mass of the gas produced at the anode is _____ g.

Part 2

21. **Define** oxidation and reduction in three different contexts: empirical (historical), theoretical (in terms of electrons), and theoretical (in terms of oxidation numbers).

22. Using a general reaction equation (A + B → C + D), label the agents and processes for any redox reaction.

23. **Define** disproportionation and provide one simple example.

24. From the information in this unit, list two or three examples of situations in which technology preceded scientific explanations.

25. Name two common reactions that occur in living and nonliving systems. For each, identify the oxidizing agent, reducing agent, and the direction of electron transfer.

26. Distinguish, in as many ways as possible, between anode and cathode. Does your answer apply equally to voltaic and electrolytic cells? Explain briefly.

27. Briefly **describe** two technological solutions to the problem of batteries "going dead."

28. **Explain** why corrosion often occurs in places where two different metals (such as copper and iron) are joined together.

29. Electrochemical cells are very important technological devices in our society. **Compare** the main differences between voltaic and electrolytic cells in terms of their purpose and the chemical reactions that occur in them.

30. **Predict** whether a spontaneous redox reaction will occur in the following situations:
(a) A copper penny is dropped into hydrochloric acid.
(b) A nickel is dropped into nitric acid.
(c) A silver earring is dropped into sulfuric acid.

31. DE While working on the development of a new electrochemical cell, a research chemist places selected Period 4 transition metal strips into aqueous solutions of their ionic compounds. She observes that the following combinations of metal and cations react spontaneously:

$V(s) + Mn^{2+}(aq) \rightarrow V^{2+}(aq) + Mn(s)$

$V^{2+}(aq) + Ti(s) \rightarrow V(s) + Ti^{2+}(aq)$

$Co^{2+}(aq) + Mn(s) \rightarrow Co(s) + Mn^{2+}(aq)$

(a) Use this information to develop a table of oxidizing and reducing agents for these metals and their ions.
(b) **Identify** the strongest oxidizing and the strongest reducing agent in your table.

32. Make a list of everything that must be balanced in a net ionic equation representing a redox reaction.

33. Write and label balanced half-reaction equations for each of the following redox reactions.
(a) $2\,Fe^{3+}(aq) + Ni(s) \rightarrow 2\,Fe^{2+}(aq) + Ni^{2+}(aq)$
(b) $Br_2(aq) + 2\,I^-(aq) \rightarrow 2\,Br^-(aq) + I_2(s)$
(c) $Pd^{2+}(aq) + Sn^{2+}(aq) \rightarrow Pd(s) + Sn^{4+}(aq)$
(d) Label each reactant in (a), (b), and (c) as an oxidizing or a reducing agent.

34. Use your knowledge of electrochemistry and some brainstorming to **describe** at least three methods for determining or approximating the position of the beryllium half-reaction in a table of half-reactions.

35. Potassium metal spontaneously reacts with water (**Figure 1**). DE

Figure 1
Potassium reacts vigorously with water.

(a) Write the half-reaction and net ionic reaction equations for this reaction.
(b) **Describe** diagnostic tests (procedure, evidence, analysis) that could be done to test for the predicted products.

36. An acidic solution of potassium dichromate is added to a sodium iodide solution. **Predict** the net ionic equation and show all of your work. State two diagnostic tests that could be used to test your reaction prediction.

37. What is the oxidation number of
(a) I in $I_2(s)$?
(b) I in $CaI_2(s)$?
(c) I in HIO(aq)?
(d) H in NH_3?
(e) H in AlH_3?
(f) O in CH_3OH?

38. **Predict** the balanced redox reaction equation by constructing and labelling oxidation and reduction half-reaction equations.
(a) $MnO_4^{2-}(aq) \rightarrow Mn^{2+}(aq) + MnO_4^-(aq)$ (acidic)
(b) $ClO^-(aq) \rightarrow ClO_2^-(aq) + Cl_2(g)$ (basic)

39. Use the oxidation number method to balance the reaction equations for the following redox reactions in acidic solutions:
(a) $MnO_4^-(aq) + H_2C_2O_4(aq) \rightarrow Mn^{2+}(aq) + CO_2(g) + H_2O(l)$
(b) $KIO_3(aq) + KI(aq) + HCl(aq) \rightarrow KCl(aq) + I_2(s) + H_2O(l)$

40. Chromium steel alloys are analyzed using a series of redox reactions.

Step 1: The alloy is initially reacted with perchloric acid, which converts the chromium metal into dichromate ions, while the perchloric acid is reduced to chlorine gas.

Step 2: The dichromate ions are then reduced to chromium(III) ions by adding an excess of iron(II) solution.

Step 3: The unreacted iron(II) is then titrated with a solution of cerium(IV) ions, which reduces them to cerium(III) ions.

Write a balanced redox equation for each step of this procedure.

41. **Predict** the mass of gold formed when a gold(III) nitrate solution reacts completely with 125 mL of 0.352 mol/L sulfurous acid (aqueous hydrogen sulfite).

42. Complete the Materials (including precautions) and Analysis of the following investigation report.

Purpose
The purpose of this exercise is to use redox stoichiometry to standardize a potassium dichromate solution.

Problem
What is the amount concentration of a potassium dichromate solution?

Evidence
Volume of iron(II) ammonium sulfate solution = 10.0 mL

Concentration of acidic $Fe(NH_4)_2(SO_4)_2(aq)$ = 0.0625 mol/L

Trial	1	2	3
final burette reading (mL)	13.4	25.5	37.6
initial burette reading (mL)	1.1	13.4	25.5

43. Electrolysis is used in the industrial production of several important elements and compounds. DE
(a) **Define** electrolysis.
(b) In an electrolytic cell, what type of half-reaction occurs at the anode? at the cathode?
(c) **Compare** the electrolysis of molten compounds with the electrolysis of aqueous solutions. State some similarities and differences.
(d) Briefly **describe** three important industrial applications of electrolysis.

44. In which of the following mixtures must an external voltage be applied to inert electrodes to observe evidence of a redox reaction?
(a) a solution of cadmium nitrate
(b) a solution of iron(III) iodide
(c) solutions of iron(III) chloride and tin(II) sulfate in separate half-cells connected by a salt bridge
(d) solutions of potassium iodide and zinc nitrate in separate half-cells

45. **Determine** the minimum potential difference that must be applied to the following electrolytic cells to cause a chemical reaction. (You do not need to write the half-cell reaction equations.)
(a) iron(II) sulfate electrolyte with inert electrodes
(b) hydrochloric acid electrolyte with silver electrodes
(c) tin(II) chloride electrolyte with tin electrodes

46. Write the equations for reactions at the cathode and anode, and the net cell reaction. **Determine** the minimum potential difference that must be applied to each of the following electrolytic cells to cause a reaction.
(a) C(s) | NaBr(aq) | C(s)
(b) Pt(s) | KOH(aq) | Pt(s)
(c) C(s) | $CoCl_2(aq)$ | C(s)

47. Volta's invention of the electric battery in 1800 led to a DE flurry of scientific research using this new technology. A few weeks after he heard about Volta's battery, William Nicholson, an English chemist, built his own battery and passed a current through slightly acidified water. With the current flowing, bubbles of colourless gases formed at each electrode. This was the first demonstration that an electric current could bring about a chemical reaction.
(a) Write equations for the cathode, anode, and net reactions that occurred in Nicholson's demonstration.
(b) Determine the minimum potential difference needed for the reaction.

48. Potassium hydroxide is obtained commercially by the DE electrolysis of aqueous potassium chloride.
(a) **Sketch** a diagram of a cell that could be used to electrolyze an aqueous solution of potassium chloride. Label electrodes, electrolyte, power supply, and the directions of the electron and ion flow.
(b) **Predict** the cathode, anode, and net reactions, and calculate the minimum potential difference for the electrolysis of aqueous potassium chloride.

49. One technological process for refining zinc metal involves DE the electrolysis of a zinc sulfate solution.
(a) Write equations for the cathode, anode, and net reactions, and calculate the minimum potential difference for the electrolysis of a zinc sulfate solution.
(b) Calculate the time required to produce 1.00 kg of zinc using a 5.00 kA current.

50. **Determine** the current required to produce 1.00 kg of aluminium per hour in a single Hall–Héroult cell for the production of aluminium.

51. The electrolysis of copper(II) sulfate using carbon electrodes is demonstrated to a chemistry class. In the demonstration, a 1.50 A current passes through 75.0 mL of 0.125 mol/L copper(II) sulfate solution. How long, in minutes, would it take to plate all of the copper from the solution?

52. Given that the typical current used in the chlor–alkali plant is 55 kA, **predict** the rate at which chlorine gas is produced (in moles per hour).

53. Battery technology is a very active area of research. One proposal that shows some promise is a vanadium redox flow cell, also known as the All Vanadium Redox Battery. **Describe** the general construction of this battery, including electrodes, electrolytes, porous boundary, and external tanks. What redox reactions occur at the electrodes within this cell? List some unique aspects of this technology and some advantages and proposed uses.

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54. Rechargeable nickel–metal hydride (NiMH) batteries have DE twice the energy density of Ni–Cd batteries and a similar operating voltage. The cells in NiMH batteries use NiO(OH)(s) as one electrode, a hydrogen-absorbing alloy as the other, and an alkaline electrolyte. In the following reduction half-equations, M indicates a hydrogen-absorbing alloy and H_{ab} indicates absorbed hydrogen.

$$NiO(OH)(s) + H_2O(l) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq) \quad E°_r = +0.49\ V$$

$$M(s) + H_2O(l) + e^- \rightarrow MH_{ab} + OH^-(aq) \quad E°_r = -0.71V$$

(a) Write balanced equations for the anode, cathode, and net reactions occurring during the cell's operation.
(b) **Determine** the cell potential.
(c) Research and list some of the technological, economic, and environmental considerations involved in evaluating the NiMH battery.

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55. Electroplating finishes are often layered. For example, DE chromium plating does not work well on a zinc base, so a layer of copper is applied to the zinc, then a layer of nickel is added, and then the top chromium layer is plated on.
(a) Propose a general design of an experiment to place a final chromium layer onto a galvanized metal. Include a labelled diagram and general plan.
(b) In any electroplating process a particular thickness of metal is desired. **Outline** the experimental variables and the type of calculations necessary to plan a particular thickness of metal plating.

56. Review the focusing questions on page 552. Using the knowledge you have gained from this unit, briefly **outline** a response to each of these questions.

Extension

57. The Alberta government is considering making carbon monoxide detectors mandatory for new residential homes as is done in Ontario. What are the sources and the health effects of carbon monoxide? Briefly **describe** how a modern CO detector works including the reaction equations and catalyst involved. Should CO detectors be required in all homes? Justify your opinion.

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58. Moli Energy of Maple Ridge, British Columbia, was the first company in the world to develop a commercial, rechargeable lithium ion cell, called a Molicel (**Figure 2**). Research the characteristics and advantages of Molicels compared with other secondary cells.

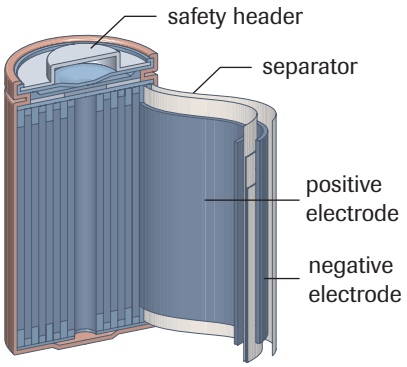

Figure 2
The Molicel is a high-energy, rechargeable lithium–ion cell in a unique, jellyroll design.